CHEMISTRY

9

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بِسۡمِ اللّٰہِ الرَّحۡمٰنِ الرَّحِیۡمِ ○

ترجمہ: ''شروع اللہ کے نام سے جو بڑا مہربان نہایت رحم والا ہے۔''

CHEMISTRY 9

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CONTENTS

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Fundamentals of Chemistry

Major Concepts

1.1 Branches of Chemistry
1.2 Basic Definitions
1.3 Chemical species
1.4 Avogadro's Number and Mole
1.5 Chemical Calculations

Time allocation	
Teaching periods	12
Assessment periods	03
Weightage	10%

Students Learning Outcomes

Students will be able to:

- Identify and provide examples of different branches of chemistry.
- Differentiate among branches of chemistry.
- Distinguish between matter and a substance.
- Define ions, molecular ions, formula units and free radicals.
- Define atomic number, atomic mass, atomic mass unit.
- Differentiate among elements, compounds and mixtures.
- Define relative atomic mass based on C-12 scale.
- Differentiate between empirical and molecular formula.
- Distinguish between atoms and ions.
- Differentiate between molecules and molecular ions.
- Distinguish between ions and free radical.
- Classify the chemical species from given examples.
- Identify the representative particles of elements and compounds.
- Relate gram atomic mass, gram molecular mass and gram formula mass to mole.
- Describe how Avogadro's number is related to a mole of any substance.
- Distinguish among the terms gram atomic mass, gram molecular mass and gram formula mass.
- Change atomic mass, molecular mass and formula mass into gram atomic mass, gram molecular mass and gram formula mass.

Introduction

The knowledge that provides understanding of this world and how it works, is science. *The branch of science which deals with the composition, structure, properties and reactions of matter is called chemistry.* It deals with every aspect of our life.

The development of science and technology has provided us a lot of facilities in daily life. Imagine the role and importance of petrochemical products, medicines and drugs, soap, detergents, paper, plastics, paints and pigments, insecticides, pesticides which all are fruit of the efforts of chemists. The development of chemical industry has also generated toxic wastes, contaminated water and polluted air around us. On the other hand, chemistry also provides knowledge and techniques to improve our health and environment and to explore and to conserve the natural resources.

In this chapter we will study about different branches of chemistry, basic definitions and concepts of chemistry.

1.1 BRANCHES OF CHEMISTRY

It is a fact that we live in the world of chemicals. We all depend upon different living organisms which require water, oxygen or carbon dioxide for their survival. Today chemistry has a wide scope in all aspects of life and is serving the humanity day and night. Chemistry is divided into following main branches: physical chemistry, organic chemistry, inorganic chemistry, biochemistry, industrial chemistry, nuclear chemistry, environmental chemistry and analytical chemistry.

1.1.1 Physical Chemistry

Physical Chemistry is defined as the *branch of chemistry that deals with the relationship between the composition and physical properties of matter along with the changes in them.* The properties such as structure of atoms or formation of molecules, behavior of gases, liquids and solids and the study of the effect of temperature or radiation on matter, all are studied under this branch.

1.1.2 Organic Chemistry

Organic Chemistry is the study of *covalent compounds of carbon and hydrogen (hydrocarbons) and their derivatives.* Organic compounds occur naturally and are also synthesized in the laboratories. Organic chemists determine the structure and properties of these naturally occurring as well as synthesized compounds. Scope of this branch covers petroleum, petrochemicals and pharmaceutical industries.

1.1.3 Inorganic Chemistry

Inorganic chemistry *deals with the study of all elements and their compounds except those of compounds of carbon and hydrogen (hydrocarbons) and their derivatives.* It has applications in every aspect of the chemical industry such as glass, cement, ceramics and metallurgy (extraction of metals from ores).

1.1.4 Biochemistry

It is *the branch of chemistry in which we study the structure, composition, and chemical reactions of substances found in living organisms.* It covers all chemical processes taking place in living organisms, such as synthesis and metabolism of biomolecules like carbohydrates, proteins and fats. Biochemistry emerged as a separate discipline when scientists began to study how living things obtain energy from food or how the fundamental biological changes occur during a disease. Examples of applications of biochemistry are in the fields of medicine, food science and agriculture etc.

1.1.5 Industrial Chemistry

The *branch of chemistry that deals with the manufacturing of chemical compounds on commercial scale, is called* **industrial chemistry.** It deals with the manufacturing of basic chemicals such as oxygen, chlorine, ammonia, caustic soda, nitric acid and sulphuric acid. These chemicals provide the raw materials for many other industries such as fertilizers, soap, textiles, agricultural products, paints and paper etc.

1.1.6 Nuclear Chemistry

Nuclear Chemistry is the *branch of chemistry that deals with the radioactivity, nuclear processes and properties*. The main concern of this branch is with the atomic energy and its uses in daily life. It also includes the study of the chemical effects resulting from the absorption of radiation within animals, plants, and other materials. It has vast applications in medical treatment (radiotherapy), preservation of food and generation of electrical power through nuclear reactors, etc.

1.1.7 Environmental Chemistry

It is *the branch of chemistry in which we study about components of the environment and the effects of human activities on the environment.* Environmental chemistry is related to other branches like biology, geology, ecology, soil, and water. The knowledge of chemical processes taking place in environment is necessary for its improvement and protection against pollution.

1.1.8 Analytical Chemistry

Analytical chemistry is *the branch of chemistry that deals with separation and analysis of a sample to identify its components.* The separation is carried out prior to qualitative and quantitative analysis. Qualitative analysis provides the identity of a substance (composition of chemical species). On the other hand quantitative analysis determines the amount of each component present in the sample. Hence, in this branch different techniques and instruments used for analysis are studied. The scope of this branch covers food, water, environmental and clinical analyses.

Test yourself 1.1

i. *In which branch of chemistry behaviour of gases and liquids is studied?*
ii. *Define biochemistry?*
iii. *Which branch of chemistry deals with preparation of paints and paper?*
iv. *In which branch of chemistry are the metabolic processes of carbohydrates and proteins studied?*
v. *Which branch of chemistry deals with atomic energy and its uses in daily life?*
vi. *Which branch of chemistry deals with the structure and properties of naturally occurring molecules?*

1.2 BASIC DEFINITIONS

Matter is simply defined as *anything that has mass and occupies space*. Our bodies as well as all the things around us are examples of matter. In chemistry we study all types of matters that can exist in any of three physical states; solid, liquid or gas.

A piece of matter in pure form is termed as a ***substance***. Every substance has a fixed composition and specific properties or characteristics. Whereas, *impure matter is called a* ***mixture***; which can be homogeneous or heterogeneous in its composition.

We know that every substance has physical as well as chemical properties. *The properties that are associated with the physical state of a matter are called* ***physical properties;*** like colour, smell, taste, hardness, shape of crystal, solubility, melting or boiling points etc. For example when ice is heated, it melts to form water. When water is further heated, it boils to give steam. In this entire process only the physical state of water changes whereas its chemical composition remains the same.

The ***chemical properties*** *depend upon the composition of the substance*. When a substance undergoes a chemical change, its composition changes and a new substances are formed. For example, decomposition of water is a chemical change as it produces hydrogen and oxygen gases. All materials are either a substance or a mixture. Figure 1.1 shows simple classification of the matter into different forms.

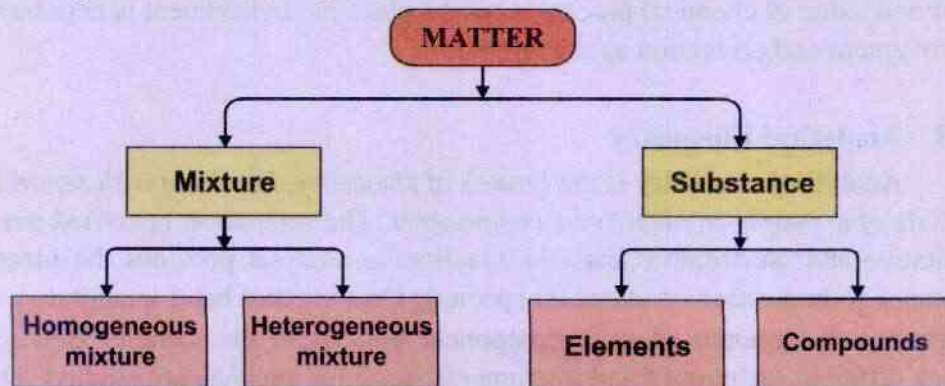



Fig. 1.1 Classification of matter

1.2.1 Elements, Compounds and Mixtures

1.2.1.1 Elements

In the early ages, only nine elements (carbon, gold, silver, tin, mercury, lead, copper, iron and sulphur) were known. At that time it was considered that elements were the substances that could not be broken down into simpler units by ordinary chemical processes. Until the end of nineteenth century sixty-three elements had been discovered. Now 118 elements have been discovered, out of which 92 are naturally occurring elements. Modern definition of **element** is that *it is a substance made up of same type of atoms, having same atomic number and it cannot be decomposed into simple substances by ordinary chemical means.* It means that each element is made up of unique type of atoms that have very specific properties.

Elements occur in nature in free or combined form. All the naturally occurring elements found in the world have different percentages in the earth's crust, oceans and atmosphere. Table 1.1. shows natural occurrence in percentage by weight of some major elements around us. It shows concentrations of these major elements found in the three main systems of our environment.

Table 1.1 Natural Occurrences by Weight % of Some Major Elements

Crust of Earth		Oceans		Atmosphere	
Oxygen	47 %	Oxygen	86 %	Nitrogen	78 %
Silicon	28 %	Hydrogen	11 %	Oxygen	21%
Aluminium	7.8 %	Chlorine	1.8 %	Argon	0.9%

Elements may be solids, liquids or gases. Majority of the elements exist as solids e.g. sodium, copper, zinc, gold etc. There are very few elements which occur in liquid state e.g. mercury and bromine. A few elements exist as gases e.g. nitrogen, oxygen, chlorine and hydrogen.

On the basis of their properties, elements are divided into metals, non-metals and metalloids. About 80 percent of the elements are metals.

Do you know?

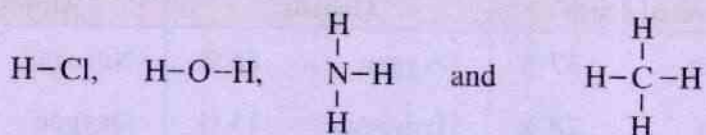
Major part of a living body is made up of water i.e. 65% to 80% by mass.
Six elements constitute about 99% of our body mass; namely: Oxygen 65 %, Carbon 18%, Hydrogen 10 %, Nitrogen 3%, Calcium 1.5% and Phosphorus 1.5%.

Potassium, Sulphur, Magnesium and Sodium constitute 0.8% of our body mass. Whereas Copper, Zinc, Fluorine, Chlorine, Iron, Cobalt and Manganese constitute only 0.2% of our body mass.

The elements are represented by **symbols**, *which are abbreviations for the name of elements.* A symbol is taken from the name of that element in English, Latin, Greek or German. If it is one-letter, it will be capital as H for Hydrogen, N for Nitrogen, and C for Carbon etc. In case of two letters symbol, only first letter is capital e.g. Ca for Calcium, Na for Sodium and Cl for Chlorine.

The unique *property of an element is* ***valency****. It is combining capacity of an element with other elements.* It depends upon the number of electrons in the outermost shell.

In simple covalent compounds, valency is the number of hydrogen atoms which combine with one atom of that element or the number of bonds formed by one atom of that element e.g. in the following compounds

$$H-Cl, \quad H-O-H, \quad \begin{array}{c} H \\ | \\ N-H \\ | \\ H \end{array} \quad \text{and} \quad \begin{array}{c} H \\ | \\ H-C-H \\ | \\ H \end{array}$$

the valency of chlorine, oxygen, nitrogen and carbon is 1, 2, 3 and 4 respectively.

In simple ionic compounds valency is the number of electrons gained or lost by an atom of an element to complete its octet. Elements having less than four electrons in their valence shell; prefer to lose the electrons to complete their octet. For example atoms of Na, Mg and Al have 1,2 and 3 electrons in their valence shells respectively. They lose these electrons to have valency of 1,2 and 3 respectively. On the other hand elements having five or more than five electrons in their valence shells, gain electrons to complete their octet. For example, N, O and Cl have 5,6 and 7 electrons in their valence shells respectively. They gain 3, 2 and 1 electrons respectively to complete their octet. Hence they show valency of 3, 2 and 1 respectively. A radical is a group of atoms that have some charge. Valencies of some common elements and radicals are shown in Table 1.2.

Table 1.2 Some Elements and Radicals with their Symbols and Common Valencies

Element / Radical	Symbol	Valency	Element / Radical	Symbol	Valency
Sodium	Na	1	Hydrogen	H	1
Potassium	K	1	Chlorine	Cl	1
Silver	Ag	1	Bromine	Br	1
Magnesium	Mg	2	Iodine	I	1
Calcium	Ca	2	Oxygen	O	2
Barium	Ba	2	Sulphur	S	2
Zinc	Zn	2	Nitrogen	N	3
Copper	Cu	1,2	Phosphorus	P	3,5
Mercury	Hg	1,2	Boron	B	3
Iron	Fe	2,3	Arsenic	As	3
Aluminium	Al	3	Carbon	C	4
Chromium	Cr	3	Carbonate	CO_3^{2-}	2
Ammonium	NH_4^+	1	Sulphate	SO_4^{2-}	2
Hydronium	H_3O^+	1	Sulphite	SO_3^{2-}	2
Hydroxide	OH^-	1	Thiosulphate	$S_2O_3^{2-}$	2
Cyanide	CN^-	1	Nitride	N^{3-}	3
Bisulphate	HSO_4^-	1	Phosphate	PO_4^{3-}	3
Bicarbonate	HCO_3^-	1			

Some elements show more than one valency, i.e., they have variable valency. For example, in ferrous sulphate ($FeSO_4$) the valency of iron is 2. In ferric sulphate ($Fe_2(SO_4)_3$) the valency of iron is 3. Generally, the Latin or Greek name for the element (e.g., Ferrum) is modified to end in 'ous' for the lower valency (e.g., Ferrous) and to end in 'ic' for the higher valency (e.g., Ferric).

1.2.1.2 Compound

Compound is a *substance made up of two or more elements chemically combined together in a fixed ratio by mass*. As a result of this combination, elements lose their own properties and produce new substances (compounds) that have entirely different properties. Compounds can't be broken down into its constituent elements by simple physical methods. For example, carbon dioxide is formed when elements of carbon and oxygen combine chemically in a fixed ratio of 12:32 or 3:8 by mass. Similarly water is a compound formed by a chemical combination between hydrogen and oxygen in a fixed ratio of 1:8 by mass.

Compounds can be classified as ionic or covalent. Ionic compounds do not exist in independent molecular form. They form a three-dimensional crystal lattice, in which each ion is surrounded by oppositely charged ions. These oppositely charged ions attract each other very strongly, as a result ionic compounds have high melting and boiling points. These compounds are represented by **formula units** e.g. NaCl, KBr, $CuSO_4$.

The covalent compounds mostly exist in molecular form. A molecule is a true representative of the covalent compound and its formula is called **molecular formula** e.g. H_2O, HCl, H_2SO_4, CH_4.

Table 1.3 Some common Compounds with their Formulae

Compound	Chemical Formula
Water	H_2O
Sodium chloride (Common salt)	NaCl
Silicon dioxide (Sand)	SiO_2
Sodium hydroxide (Caustic Soda)	NaOH
Sodium carbonate (Washing Soda)	$Na_2CO_3.10H_2O$
Calcium oxide (Quick Lime)	CaO
Calcium carbonate (Lime Stone)	$CaCO_3$
Sugar	$C_{12}H_{22}O_{11}$
Sulphuric acid	H_2SO_4
Ammonia	NH_3

Remember

Always use:

- Standard symbols of elements
- Chemical formulas of compounds
- Proper abbreviations of scientific terms
- Standard values and SI units for constants

1.2.1.3 Mixture

When two or more elements or compounds mix-up physically without any fixed ratio, they form a ***mixture.*** On mixing up, the component substances retain their own chemical identities and properties. The mixture can be separated into parent components by physical methods such as distillation, filtration, evaporation, crystallisation or magnetization. *Mixtures that have uniform composition throughout are called* ***homogeneous mixtures*** e.g. air, gasoline, ice cream. Whereas, ***heterogeneous mixtures*** *are those in which composition is not uniform throughout* e.g. soil, rock and wood.

Do you know?

Air *is a mixture of nitrogen, oxygen, carbon dioxide, noble gases and water vapours.*

Soil *is a mixture of sand, clay, mineral salts, water and air.*

Milk *is a mixture of water, sugar, fat, proteins, mineral salts and vitamins.*

Brass *is a mixture of copper and zinc metals.*

Table 1.4 Difference between a Compound and a Mixture

	Compound	Mixture
i.	It is formed by a chemical combination of atoms of the elements.	Mixture is formed by the simple mixing up of the substances.
ii.	The constituents lose their identity and form a new substance having entirely different properties from them.	Mixture shows the properties of the constituents.
iii.	Compounds always have fixed composition by mass.	Mixtures do not have fixed composition.
iv.	The components cannot be separated by physical means.	The components can be separated by simple physical methods.
v.	Every compound is represented by a chemical formula.	It consists of two or more components and does not have any chemical formula.
vi.	Compounds have homogeneous composition.	They may be homogeneous or heterogeneous in composition
vii.	Compounds have sharp and fixed melting points.	Mixtures do not have sharp and fixed melting points.

Test yourself 1.2

i. Can you identify mixture, element or compound out of the followings:
Coca cola, petroleum, sugar, table salt, blood, gun powder, urine, aluminium, silicon, tin, lime and ice cream.

ii. How can you justify that air is a homogenous mixture. Identify substances present in it.

iii. Name the elements represented by the following symbols:
Hg, Au, Fe, Ni, Co, W, Sn, Na, Ba, Br, Bi.

iv. Name a solid, a liquid and a gaseous element that exists at the room temperature.

v. What elements do the following compounds contain?
Sugar, common salt, lime water and chalk.

1.2.1 Atomic Number and Mass Number

The **atomic number** *of an element is equal to the number of protons present in the nucleus of its atoms*. It is represented by symbol '**Z**'. As all atoms of an element have the same number of protons in their nuclei, they have the same atomic number.

Hence each element has a specific atomic number termed as its identification number. For example, all hydrogen atoms have 1 proton, their atomic number **Z**=1. All atoms in carbon have 6 protons, their atomic number **Z**=6. Similarly, in oxygen all atoms have 8 protons having atomic number **Z**=8 and sulphur having 16 protons shows atomic number **Z** = 16.

The **mass number** *is the sum of number of protons and neutrons present in the nucleus of an atom.* It is represented by symbol '**A**'.

It is calculated as $A=Z+n$ *where* n is the number of neutrons.

Each proton and neutron has 1amu mass. For example hydrogen atom has one proton and no neutron in its nucleus, its mass number A=1+0 =1. Carbon atom has 6 protons and 6 neutrons, hence its mass number A=12. Atomic numbers and mass numbers of a few elements are given in Table 1.5

Table 1.5 Some Elements along with their Atomic and Mass Numbers

Element	Number of Protons	Number of Neutrons	Atomic Number Z	Mass Number A
Hydrogen	1	0	1	1
Carbon	6	6	6	12
Nitrogen	7	7	7	14
Oxygen	8	8	8	16
Fluorine	9	10	9	19
Sodium	11	12	11	23
Magnesium	12	12	12	24
Potassium	19	20	19	39
Calcium	20	20	20	40

Example 1.1

How many protons and neutrons are there in an atom having A = 238 and Z = 92.

Solution:

First of all develop data from the given statement of the example and then solve the it with the help of data.

Data

A=238

Z=92

Number of protons ?

Number of neutrons?

Number of protons = Z = 92

Number of Neutrons $= A - Z$
$= 238 - 92$
$= 146$

1.2.3 Relative Atomic Mass and Atomic Mass Unit

As we know that the mass of an atom is too small to be determined practically. However, certain instruments enable us to determine the ratio of the atomic masses of various elements to that of carbon-12 atoms. This ratio is known as the relative atomic mass of the element. The **relative atomic mass** *of an element is the average mass of an atoms of that element as compared to 1/12th (one-twelfth) the mass of an atom of carbon-12 isotope (an element having different mass number but same atomic number).* Based on carbon-12 standard, the mass of an atom of carbon is 12 and 1/12th of it comes to be one. When we compare atomic masses of other elements with carbon-12 atoms, they are expressed as relative atomic masses of those elements. The unit for relative atomic masses is called **atomic mass unit,** with symbol ***'amu'***. One atomic mass unit *is 1/12th the mass of one atom of carbon-12th*. When this atomic mass unit is expressed in ***grams,*** it is:

$$1\ \textbf{amu} = 1.66 \times 10^{-24} g$$

For example:

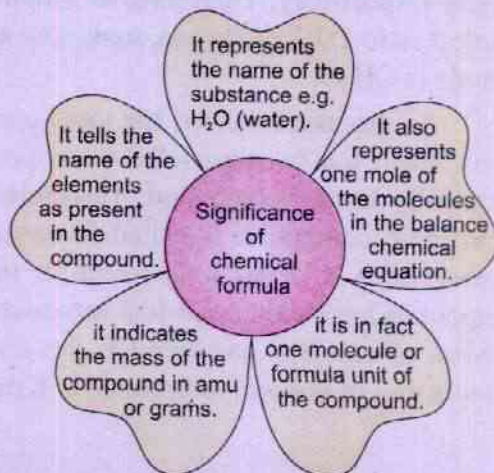
Mass of a proton = 1.0073 *amu* or 1.672×10^{-24} *g*
Mass of a neutron = 1.0087 *amu* or 1.674×10^{-24} *g*
Mass of an electron = 5.486×10^{-4} *amu.* or 9.106×10^{-28} *g*

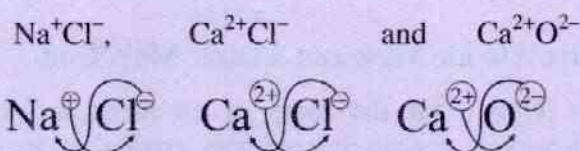
Test yourself 1.3

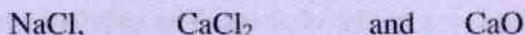
i) How many amu 1 g of a substance has?
ii) Is atomic mass unit a SI unit of an atomic mass?
iii) What is the relationship between atomic number and atomic mass?
iv) Define relative atomic mass.
v) Why atomic mass of an atom is defined as relative atomic mass?

1.2.4 How to write a Chemical Formula

Compounds are represented by **chemical formulae** as elements are represented by symbols. Chemical formulae of compounds are written keeping the following steps in consideration.

i. Symbols of two elements are written side-by-side, in the order of positive ion first and negative ion later.
ii. The valency of each ion is written on the right top corner of its symbol, e.g. Na^+, Ca^{2+}, Cl^- and O^{2-}.

iii. This valency of each ion is brought to the lower right corner of other ion by 'cross-exchange' method, e.g.

Na^+Cl^-, $Ca^{2+}Cl^-$ and $Ca^{2+}O^{2-}$

They are written as:

$NaCl$, $CaCl_2$ and CaO

iv. If the valencies are same, they are offset and are not written in the chemical formula. But if they are different, they are indicated as such at the same position, e.g. in case of sodium chloride both the valencies are offset and formula is written as NaCl, whereas, calcium chloride is represented by formula $CaCl_2$.

v. If an ion is a combination of two or more atoms which is called radical, bearing a net charge on it, e.g. SO_4^{2-} (sulphate) and PO_4^{3-} (phosphate), then the net charge represents the valency of the radical. The chemical formula of such compounds is written as explained in (iii) and (iv); writing the negative radical within the parenthesis. For example, chemical formula of aluminium sulphate is written as $Al_2(SO_4)_3$ and that of calcium phosphate as $Ca_3(PO_4)_2$.

1.2.4.1 Empirical formula

Chemical formulae are of two types. The simplest type of formula is **empirical formula.** It is the *simplest whole number ratio of atoms present in a compound.* The empirical formula of a compound is determined by knowing the percentage composition of a compound. However, here we will explain it with simple examples.

The covalent compound silica (sand) has simplest ratio of 1:2 of silicon and oxygen respectively. Therefore, its empirical formula is SiO_2. Similarly, glucose has simplest ratio 1:2:1 of carbon, hydrogen and oxygen respectively. Hence its empirical formula is CH_2O.

As discussed earlier, the ionic compounds exist in three dimensional network forms. Each ion is surrounded by oppositely charged ions in such a way to form electrically neutral compound. Therefore, the simplest unit taken as a representative of an ionic compound is called **formula unit**. It is *defined as the simplest whole number ratio of ions, as present in the ionic compound.* In other words, ionic compounds have only empirical formulae. For example, formula unit of common salt consists of one Na^+ and one Cl^- ion and its empirical formula is NaCl. Similarly, formula unit of potassium bromide is KBr, which is also its empirical formula.

1.2.4.2 Molecular Formula

Molecules are formed by the combination of atoms. These molecules are represented by **molecular formulae** that show *actual number of atoms of each element present in a molecule of that compound.* Molecular formula is derived from empirical formula by the following relationship:

$$\text{Molecular formula} = (\text{Empirical formula})_n$$

Where n is 1,2,3 and so on.

For example, molecular formula of benzene is C_6H_6 which is derived from the empirical formula CH where the value of n is 6.

The molecular formula of a compound may be same or a multiple of the empirical formula. A few compounds having different empirical and molecular formulae are shown in Table 1.6.

Table 1.6 Some Compounds with their Empirical and Molecular Formulae

Compound	Empirical formula	Molecular formula
Hydrogen peroxide	HO	H_2O_2
Benzene	CH	C_6H_6
Glucose	CH_2O	$C_6H_{12}O_6$

Some compounds may have same empirical and molecular formula e.g. water (H_2O), hydrochloric acid (HCl), etc.

1.2.5 Molecular Mass and Formula Mass

The sum of atomic masses of all the atoms present in one molecule of a molecular substance, is its **molecular mass.** For example molecular mass of chlorine (cl_2) is 71.0 *amu,* of water (H_2O) is 18 *amu* and that of carbon oxide (CO_2) is 44 *amu.*

Example 1.2

Calculate the molecular mass of Nitric acid, HNO_3.

Solution

Atomic mass of H = 1 *amu*
Atomic mass of N = 14 *amu*
Atomic mass of O = 16 amu
Molecular formula = HNO_3

$$\begin{aligned}\text{Molecular mass} &= 1\,(\text{At. mass H}) + 1\,(\text{At. mass N}) + 3\,(\text{At. mass O}) \\ &= 1 + 14 + 3(16) \\ &= 1 + 14 + 48 \\ &= 63\ amu\end{aligned}$$

Some ionic compounds that form three-dimensional solid crystals, are represented by their formula units. **Formula mass** in such cases *is the sum of atomic*

masses of all the atoms present in one formula unit of a substance. For example formula mass of sodium chloride is 58.5 *amu* and that of $CaCO_3$ is 100 *amu.*

Example 1.3

Calculate the formula mass of Potassium sulphate K_2SO_4

Solution

Atomic mass of K	=	39 *amu*
Atomic mass of S	=	32 *amu*
Atomic mass of O	=	16 *amu*
Formula unit	=	K_2SO_4
Formula mass of K_2SO_4	=	2(39) + 1(32) + 4(16)
	=	78 + 32 + 64
	=	174 *amu*

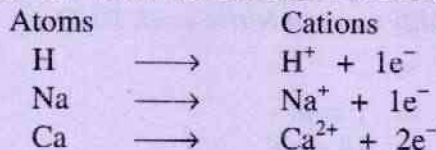
Test yourself 1.4

i. What is the relationship beeween empirical formula and formula unit?
ii. How can you differentiate between molecular formula and empirical formula?
iii. Identify the followings as molecular ionic carped:
H_2O_2, CH_4, $C_6H_{12}O_6$, $C_{12}H_{22}O_{11}$, $BaCO_3$, KBr
iv. What is empirical formula of acetic acid (CH_3COOH)? Find out its molecular mass.
v. Calculate the formula masses of:
Na_2SO_4, $ZnSO_4$ and $CuCO_3$.

1.3 CHEMICAL SPECIES

1.3.1 Ions (Cations and Anions), Molecular Ions and Free Radicals

Ion *is an atom or group of atoms having a charge on it.* The charge may be positive or negative. There are two types of ions i.e. cations and anions. *An atom or group of atoms having positive charge on it is called* **cation.** The cations are formed when an atom loses electrons from its outermost shells. For example, Na^+, K^+ are cations. The following equations show the formation of cations from atoms.

Atoms		Cations
H	⟶	$H^+ + 1e^-$
Na	⟶	$Na^+ + 1e^-$
Ca	⟶	$Ca^{2+} + 2e^-$

An atom or a group of atoms that has a negative charge on it, is called **anion.** Anion is formed by the gain or addition of electrons to an atom. For example, Cl^- and O^{2-}. Following examples show the formation of an anion by addition of electrons to an atom.

Atoms				Anions
Cl	+	$1e^-$	⟶	Cl^-
O	+	$2e^-$	⟶	O^{2-}

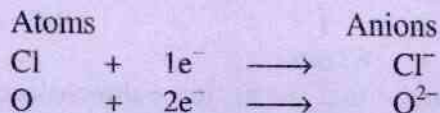



Table 1.7 Difference between Atoms and Ions

	Atom	Ion
i.	It is the smallest particle of an element.	It is the smallest unit of an ionic compound.
ii.	It can or cannot exist independently and can take part in a chemical reaction.	It cannot exist independently and is surrounded by oppositely charged ions.
iii.	It is electrically neutral.	It has a net charge (either negative or positive) on it.

1.3.1.1 Molecular Ion

When a molecule loses or gains an electron, it forms a **molecular ion**. Hence, molecular ion or radical is a species having positive or negative charge on it. Like other ions they can be cationic molecular ions (if they carry positive charge) or anionic molecular ions (if they carry negative charge). Cationic molecular ions are more abundant than anionic molecular ions. For example CH_4^+, He^+, N_2^+. When gases are bombarded with high-energy electrons in a discharge tube, they ionize to give molecular ions. Table 1.8 shows some differences between molecule and molecular ion.

Table 1.8 Difference between Molecule and Molecular Ion

	Molecule	Molecular Ion
i.	It is the smallest particle of an element or compound which can exist independently and shows all the properties of that compound.	It is formed by gain or loss of electrons by a molecule.
ii.	It is always neutral.	It can have negative or positive charge
iii.	It is formed by the combination of atoms.	It is formed by the ionization of a molecule.
iv.	It is a stable unit.	It is a reactive specie.

1.3.1.2 Free Radicals

Free radicals *are atoms or group of atoms possessing odd number of (unpaired) electrons*. It is represented by putting a dot over the symbol of an element e.g. $H^\bullet$, $Cl^\bullet$, $H_3C^\bullet$. Free radicals are generated by the homolytic (equal) breakage of the bond between two atoms when they absorb heat or light energy. A free radical is extremely reactive species as it has the tendency to complete its octet. Table 1.9 shows some of the differences between ions and free radicals.

Do you know?

Most of the universe exists in the form of plasma, the fourth state of matter. Both the cationic and anionic molecular ions are present in it.

Molecules		Free radicals
Cl_2	$\xrightarrow{\text{sunlight}}$	$2Cl^{\bullet}$
CH_4	$\xrightarrow{\text{sunlight}}$	$H_3C^{\bullet} + H^{\bullet}$

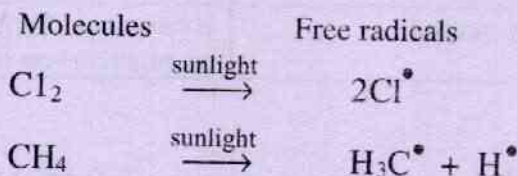



Table 1.9 Difference between Ions and Free Radicals

	Ions	Free Radicals
i.	These are the atoms which bear some charge.	These are the atoms that have odd number of electrons.
ii.	They exist in solution or in crystal lattice.	They can exist in solutions as well in air.
iii.	Their formation is not affected by the presence of light.	They may form in the presence of light.

1.3.2 Types of Molecules

A **molecule** *is formed by the chemical combinations of atoms. It is the smallest unit of a substance*. It shows all the properties of the substance and can exist independently. There are different types of molecules depending upon the number and types of atoms combining. A few types are discussed here.

A molecule consisting of only one atom is called **monoatomic molecule**. For example, the inert gases helium, neon and argon all exist independently in atomic form and they are called monoatomic molecules.

If a molecule consists of two atoms, it is called ***diatomic molecule***. For example: hydrogen (H_2), oxygen (O_2), chlorine (Cl_2) and hydrogen chloride (HCl).

If it consists of three atoms, it is called ***triatomic molecule.*** For example H_2O and CO_2. If a molecule consists of many atoms, it is called polyatomic For example: methane (CH_4), sulphuric acid (H_2SO_4) and glucose ($C_6H_{12}O_6$).

A Molecule containing same type of atoms, is called ***homoatomic molecule***. For example: hydrogen (H_2), ozone (O_3), sulphur (S_8) and phosphorus (P_4) are the examples of molecules formed by the same type of atoms. *When a molecule consists of different kinds of atoms, it is called* **heteroatomic** molecule. For example: CO_2, H_2O and NH_3.

Test yourself 1.5

i. *Identify the followings as diatomic, triatomic or polyatomic molecules*
H_2SO_4, H_2, CO_2, HCl, CO, C_6H_6, H_2O

ii. *Identify among the followings as cation, anion, free radical, molecular ion or molecule:*
Na^+, $Br^{\bullet}$, N_2^+, N_2, Cl_2, CO_3^{2-}, H, O_2, O^{2-}

1.4 GRAM ATOMIC MASS, GRAM MOLECULAR MASS AND GRAM FORMULA MASS

As we know that all substances are made up of atoms, molecules or formula units. Their masses are referred to as atomic mass, molecular mass and formula mass respectively and are expressed in amu. But these masses can also be expressed in other units as well. When masses are expressed in grams, they are termed as following:

i. Gram atomic mass
ii. Gram molecular mass
iii. Gram formula mass

1.4.1 Gram Atomic Mass

The atomic mass of an element expressed in grams is called ***gram atomic mass*** *or* ***gram atom.*** It is also called a *mole.*

1 gram atom of hydrogen = 1.008 *g* = 1 mol of hydrogen atoms
1 gram atom of carbon = 12.0 *g* = 1 mol of carbon

It means that 1 gram atom of different elements has different masses.

1.4.2 Gram Molecular Mass

The molecular mass of an element or a compound expressed in grams is called ***gram molecular mass*** *or* ***gram molecule.*** It is also called a *mole.*

1 gram molecule of of H_2 = 2.0 *g* = 1 mol of hydrogen molecules
1 gram molecule of H_2O = 18.0 *g* = 1 mol of water
1 gram molecule of H_2SO_4 = 98.0 *g* = 1 mol of sulphuric acid

1.4.3 Gram Formula Mass

The formula mass of an ionic compound expressed in grams is called ***gram formula mass*** *or* ***gram formula.*** This is also called *a mole*.. For

1 gram formula of NaCl = 58.5 g = 1 mol of sodium chloride
1 gram formula of $CaCO_3$ = 100 g = 1 mol of calcium carbonate

1.5 AVOGADRO'S NUMBER AND MOLE

1.5.1 Avogadro's Number

In chemistry we deal with substances which are composed of atoms, molecules or formula units. The counting of these particles is not possible for the chemists. The concept of Avogadro's number facilitated the counting of particles

contained in the given mass of a substance. **Avogadro's Number** is a *collection of 6.02×10^{23} particles.* It is represented by symbol 'N_A'. Hence, the 6.02×10^{23} number of atoms, molecules or formula units is called Avogadro's number that is equivalent to one '*mole*' of respective substance. In simple words 6.02×10^{23} particles are equal to one mole as twelve eggs are equal to one dozen.

To understand **the relationship between the Avogadro's number and the mole** of a substance let us consider a few examples.

i. 6.02×10^{23} atoms of carbon are equivalent to one mole of carbon.

ii. 6.02×10^{23} molecules of H_2O are equivalent to one mole of water.

iii. 6.02×10^{23} formula units of NaCl are equivalent to one mole of sodium chloride.

Amaedo Avogadro (1776-1856) was an Italian scholar. He is famous for molecular theory commonly known as Avogadro's law. In tribute to him, the number of particles (atoms, molecules, ions) in mole of a substance 6.02×10^{23} is known as the Avogadro's constant.

Thus, 6.02×10^{23} atoms of elements or 6.02×10^{23} molecules of molecular substance or 6.02×10^{23} formula units of ionic compounds are equivalent to 1 mole.

For further explanation about number of atoms in molecular compounds or number of ions in ionic compounds let us discuss two examples:

i. One molecule of water is made up of 2 atoms of hydrogen and 1 atom of oxygen, hence $2 \times 6.02 \times 10^{23}$ atoms of hydrogen and 6.02×10^{23} atoms of oxygen constitute one mole of water.

ii. One formula unit of sodium chloride consists of one sodium ion and one chloride ion. So there are 6.02×10^{23} number of Na^+ ions and 6.02×10^{23} Cl^- ions in one mole of sodium chloride. Thus, the total number of ions in 1 mole of NaCl is 12.04×10^{23} or 1.204×10^{24}.

1.5.2 Mole (Chemist secret unit)

A ***mole*** *is defined as the amount(mass) of a substance that contains 6.02×10^{23} number of particles (atoms, molecules or formula units).* It establishes a link between mass of a substance and number of particles as shown in summary of molar calculations. It is abbreviated as '*mol*'.

You know that a substance may be an element or compound (molecular or ionic). Mass of a substance is either one of the following: atomic mass, molecular mass or formula mass. These masses are expressed in atomic mass units (amu). But when these masses are expressed in grams, they are called as molar masses.

Scientists have agreed that Avogadro's number of particles are present in one molar mass of a substance. *Thus, quantitative definition of* ***mole*** *is the atomic mass, molecular mass or formula mass of a substance expressed in grams is called mole.*

For example:

Atomic mass of carbon expressed as 12 *g*	= 1 mol of carbon
Molecular mass of H_2O expressed as 18 *g*	= 1 mol of water
Molecular mass of H_2SO_4 expressed as 98 *g*	= 1 mol of H_2SO_4
Formula mass of NaCl expressed as 58.5 g	= 1 mol of NaCl

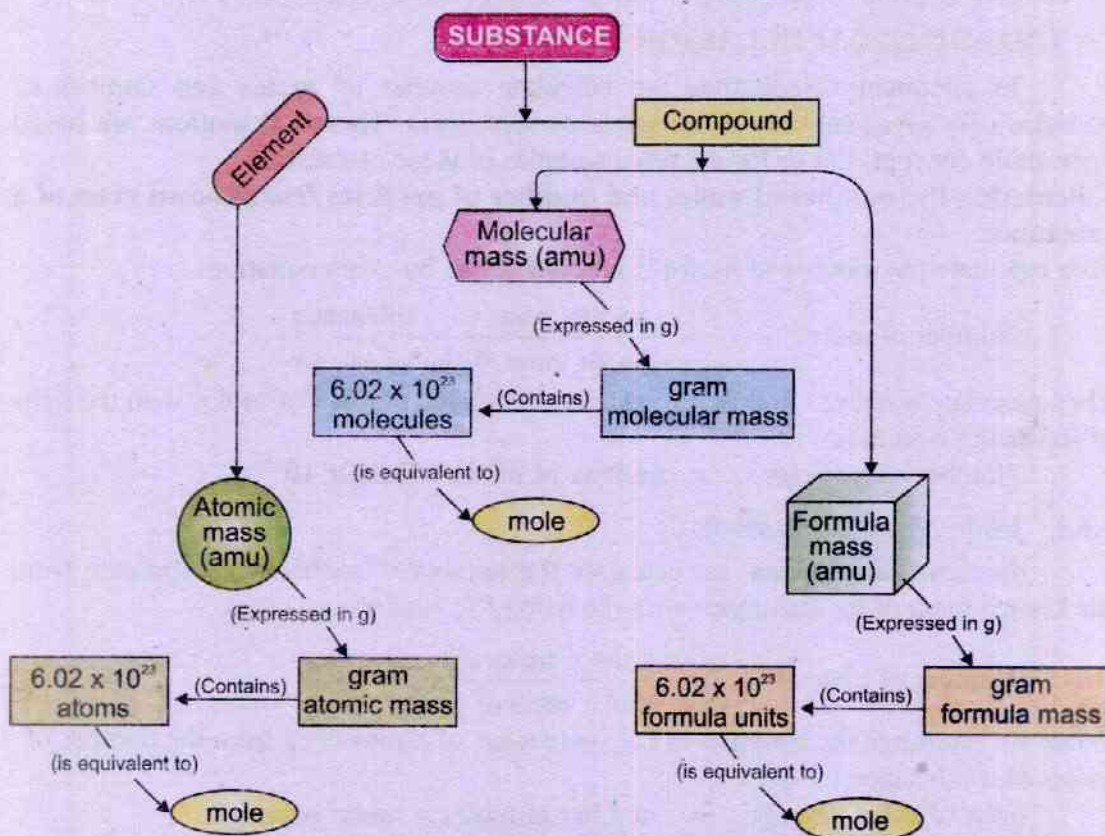
Thus, **the relationship between mole and mass** can be expressed as:

$$\text{Number of moles} = \frac{\text{known mass of a substance}}{\text{molar mass of the substance}}$$

Or,

$$\text{Mass of substance (g)} = \text{number of moles} \times \text{molar mass}$$

A detailed relationship between a substance and a mole through molar mass and number of particles is presented here.

Summary showing a relationship between a substance and a mole.

Test yourself 1.6

i. Which term is used to represent the mass of 1 mole of molecules of a substance?
ii. How many atoms are present in one gram atomic mass of a substance?
iii. Explain the relationship between mass and mole of a substance.
iv. Find out the mass of 3 moles of oxygen atoms.
v. How many molecules of water are present in half mole of water?

Example 1.4

Calculate the gram molecule (number of moles) in 40 g of H_3PO_4.

Solution

Given mass of H_3PO_4 = 40 g

Molecular mass of H_3PO_4 = 98 g mol^{-1}

Putting these values in equation

$$\text{Number of gram molecule (mol)} = \frac{\text{mass of a substance}}{\text{molar mass of the substance}}$$

$$= \frac{40}{98} = 0.408$$

Therefore, 40 grams will contain 0.408 gram molecule (mol) of H_3PO_4.

1.6 CHEMICAL CALCULATIONS

In chemical calculations we calculate number of moles and number of particles of a given mass of a substance or vice versa. These calculations are based upon mole concept. Let us have a few examples of these calculations.

Calculating the number of moles and number of particles from known mass of a substance.

First calculate the number of moles from given mass by using equation

$$\text{Number of moles} = \frac{\text{known mass of a substance}}{\text{molar mass of the substance}}$$

Then calculate number of particles from the calculated number of moles with the help of following equation:

$$\text{Number of particles} = \text{number of moles} \times 6.02 \times 10^{23}$$

1.6.1 Mole-Mass Calculations

In these calculations, we calculate the number of moles of a substance from the known mass of the substance with the help of following equation:

$$\text{Number of moles} = \frac{\text{known mass of a substance}}{\text{molar mass of the substance}}$$

When we rearrange the equation to calculate mass of a substance from the number of moles of a substance we get,

$$\text{mass of substance (g)} = \text{number of moles} \times \text{molar mass(g)}$$

Example 1.5

You have a piece of coal (carbon) weighing 9.0 gram. Calculate the number of moles of coal in the given mass.

Solution

The mass is converted to the number of moles by the equation:

$$\text{Number of moles} = \frac{\text{known mass of a substance}}{\text{molar mass of the substance}}$$

$$= \frac{9.0}{12} = 0.75$$

So, 9.0 g of coal is equivalent to 0.75 mol.

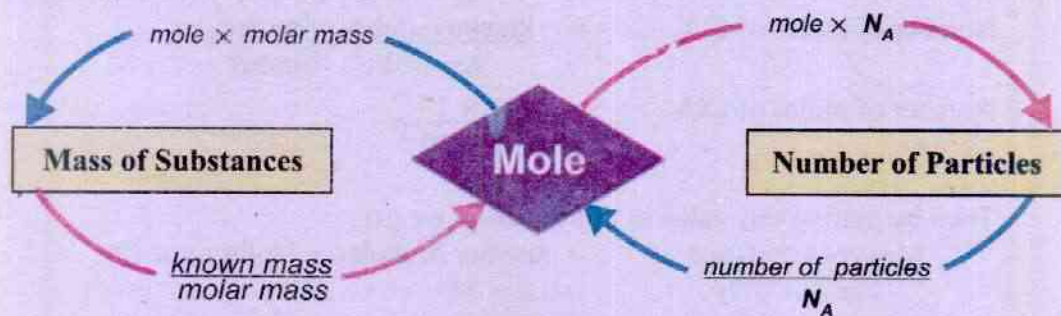
1.6.2 Mole-Particle Calculations

In these calculations we can calculate the number of *moles* of a substance from the given number of particles. (These particles are the atoms, molecules or formula units).

$$\text{Number of moles} = \frac{\text{given number of particles}}{\text{Avogadro's number}} = \frac{\text{given number of particles}}{6.02 \times 10^{23}}$$

On rearranging above equation we get,

$$\text{Number of particles} = \text{number of moles} \times 6.02 \times 10^{23}$$

Summary of Molar Calculations:

Remember

- Never calculate the number of particles from mass of the substance or vice versa. Always make calculations through moles.
- For calculations of the number of atoms in molecular compounds and the number of ions in ionic compounds; first calculate the number of molecules or formula units and then calculate the number of atoms or ions.

Example 1.6

Calculate the number of moles, number of molecules and number of atoms present in 6 grams of water.

Solution

The known mass of water $= 6\ g$

Molar mass of H_2O $= 18\ g$

Number of moles of water $= \dfrac{\text{mass of a substance}}{\text{molar mass of the substance}}$

$= 6/18$

$= 0.33\ mol$

Number of molecules = number of moles × Avogadro's number

$= 0.33 \times 6.02 \times 10^{23}$

$= 1.98 \times 10^{23}$ molecules

The number of molecules contained in 6 grams of water are 1.98×10^{23}

As we know 1 molecule of water consists of 3 atoms, therefore:

Number of atoms $= 3 \times 1.98 \times 10^{23}$

$= 5.94 \times 10^{23}$

Example 1.7

There are 3.01×10^{23} molecules of CO_2 present in a container. Calculate the number of moles and its mass.

Solution

We can calculate the number of molecules of CO_2 by putting the values in equation

Number of moles of CO_2 $= \dfrac{\text{known number of molecules}}{\text{Avogadro's Number}}$

Number of moles of CO_2 $= \dfrac{3.01 \times 10^{23}}{6.02 \times 10^{23}}$

$= 0.5$ mol

Then by putting this value in this equation we get

Mass of substance = number of moles × molar mass (g)

Mass of CO_2 $= 0.5 \times 44$

$= 22g$

Test yourself 1.7

i. How many atoms of sodium are present in 3 moles of sodium and what is the mass of it?
ii. How many atoms are in 1 amu and 1 g of hydrogen (H)?
iii. How many atoms are present in 16 g of O and 8g of S?
iv. Is the mass of 1 mole of O and 1 mole of S same?
v. What do you mean by 1 atom of C and 1 gram atom of C?
vi. If 16 g of oxygen contains 1 mole of oxygen atoms calculate the mass of one atom of oxygen in grams.
vii. How many times is 1 mole of oxygen atom heavier than 1 mole of hydrogen atom?
viii. Why does 10 g nitrogen gas contain the same number of molecules as 10 g of carbon monoxide?

THE MOLECULARITY OF THE PHYSICAL WORLD.

The nature of the physical world as perceived through men's senses has been investigated in depth. The biggest lesson we learnt in 20^{th} century is that Chemistry has become central science. It leads to the discovery of every chemical reaction in any living and non-living thing based on formation of "molecules". A reaction in the smallest living organism or in the most developed species like man, always takes place through the process of molecule formation. Hence it provides basis of "molecularity" of the physical world.

CORPUSCULAR NATURE OF MATTER.

In 1924 de Broglie put forward the theory of dual nature of matter i.e. matter has both the properties of particles as well as waves. He explained the background of two ideas. He advocated that these two systems could not remain detached from each other. By mathematical evidences he proved that every moving object is attached with waves and every wave has corpuscular nature as well. It formulated a basis to understand corpuscular nature of matter.

THE WORKS OF DIFFERENT SCIENTISTS AT THE SAME TIME HANDICAP OR PROMOTE THE GROWTH OF SCIENCE.

Over the course of human history, people have developed many interconnected and validated ideas about the physical, biological, psychological, and social worlds. Those ideas have enabled successive generations to achieve an increasingly comprehensive and reliable understanding of the human species and its environment. The means used to develop these ideas are particular ways of observing, thinking, experimenting and validating. These ways represent a fundamental aspect of the nature of science and reflect how science tends to differ from other modes of knowing. It is the union of science, mathematics and technology that forms the scientific endeavor and that makes it so successful. Although,each of these human enterprises has a character and history of its own, each is dependent on and reinforces the others.

MOLE - A QUANTITY

A computer counting with a speed of 10 million atoms a second would take 2 billion years to count one mole of atoms.

If one mole of marbles were spread over the surface of the Earth, our planet would be covered by a 3 miles thick layer of marbles.

A glass of water, which contains about 10 moles of water, has more water molecules than the grains of sand in the Sahara desert.

Key Points

- Chemistry is study of composition and properties of matter. It has different branches.
- Substances are classified into elements and compounds.
- Elements consist of only one type of atoms.
- Compounds are formed by chemical combination of atoms of the elements in a fixed ratio.
- Mixtures are formed by mixing up elements or compounds in any ratio. They are classified as homogeneous and heterogeneous mixtures.

- Each atom of an element has a specific atomic number (Z) and a mass number or atomic mass (A).
- Atomic mass of an atom is measured relative to a standard mass of C-12.
- Relative atomic mass of an element is the mass of an element compared with 1/12 mass of an atom of C-12 isotope.
- Atomic mass unit is 1/12 of the mass of one atom of C-12. 1amu = 1.66×10^{-24} g.
- Empirical formula is the simplest type of chemical formula, which shows the relative number of atoms of each element in a compound.
- Molecular formula gives the actual number of atoms of each element in a molecule.
- Formula mass is the sum of atomic masses of all the atoms in one formula unit of a substance.
- An atom or group of atoms having a charge on it is called an ion. If it has positive charge it is called a cation and if it has negative charge it is called an anion.
- There are different types of molecules: monoatomic, polyatomic, homoatomic and heteroatomic.
- The number of particles in one mole of a substance is called Avogadro's number. The value of this number is 6.02×10^{23}. It is represented as N_A.
- The amount of a substance having 6.02×10^{23} particles is called a mole. The quantitative definition of mole is atomic mass, molecular mass or formula mass expressed in grams.

EXERCISE

Multiple Choice Questions

Put a (✓) on the correct answer

1. **Industrial chemistry deals with the manufacturing of compounds:**
   (a) in the laboratory (b) on micro scale
   (c) on commercial scale (d) on economic scale
2. **Which one of the following compounds can be separated by physical means?**
   (a) mixture (b) element (c) compound (d) radical
3. **The most abundant element occurring in the oceans is:**
   (a) oxygen (b) hydrogen (c) nitrogen (d) silicon
4. **Which one of the following elements is found in most abundance in the Earth's crust?**
   (a) oxygen (b) aluminium (c) silicon (d) iron
5. **The third abundant gas found in the Earth's atmosphere is:**
   (a) carbon monoxide (b) oxygen. (c) nitrogen (d) argon
6. **One amu (atomic mass unit) is equivalent to:**
   (a) 1.66×10^{-24} mg (b) 1.66×10^{-24}g
   (c) 1.66×10^{-24}kg (d) 1.66×10^{-23}g

7. **Which one of the following molecule is not tri-atomic?**
(a) H_2 (b) O_3 (c) H_2O (d) CO_2
8. **The mass of one molecule of water is:**
(a)18 amu (b) 18 g (c) 18 mg (d) 18 kg
9. **The molar mass of H_2SO_4 is:**
(a)98 g (b) 98 amu (c) 9.8 g (d) 9.8 amu
10. **Which one of the following is a molecular mass of O_2 in amu?**
(a) 32 amu (b) 53.12×10^{-24} amu
(c) 1.92×10^{-25} amu (d) 192.64×10^{-25} amu
11. **How many number of moles are equivalent to 8 grams of CO_2?**
(a) 0.15 (b) 0.18 (c) 0.21 (d) 0.24
12. **In which one of the following pairs has the same number of ions?**
(a) 1 mole of NaCl and 1 mole of $MgCl_2$
(b) 1/2 mole of NaCl and 1/2 mole of $MgCl_2$
(c) 1/2 mole of NaCl and 1/3 mole of $MgCl_2$
(d) 1/3 mole of NaCl and 1/2 mole of $MgCl_2$
13. **Which one of the following pairs has the same mass?**
(a) 1 mole of CO and 1 mole of N_2
(b) 1 mole of CO and 1 mole of CO_2
(c) 1 mole of O_2 and 1 mole of N_2
(d) 1 mole of O_2 and 1 mole of CO_2

Short answer questions.

1. Define industrial chemistry and analytical chemistry.
2. How can you differentiate between organic and inorganic chemistry?
3. Give the scope of biochemistry.
4. How does homogeneous mixture differ from heterogeneous mixture?
5. What is the relative atomic mass? How is it related to gram?
6. Define empirical formula with an example.
7. State three reasons why do you think air is a mixture and water a compound?
8. Explain why are hydrogen and oxygen considered elements whereas water as a compound.
9. What is the significance of the symbol of an element?
10. State the reasons: soft drink is a mixture and water is a compound.
11. Classify the following into element, compound or mixture:
i. He and H_2 ii. CO and Co iii. Water and milk
iv. Gold and brass v. Iron and steel
12. Define atomic mass unit. Why is it needed?

13. State the nature and name of the substance formed by combining the following:
    i. Zinc + Copper ii. Water + Sugar
    iii. Aluminium + Sulphur iv. Iron + Chromium + Nickel
14. Differentiate between molecular mass and formula mass, which of the followings have molecular formula?
    H_2O, NaCl, KI, H_2SO_4
15. Which one has more atoms: 10 g of Al or 10 g of Fe?
16. Which one has more molecules: 9 g of water or 9 g of sugar ($C_{12}H_{22}O_{11}$)?
17. Which one has more formula units: 1 g of NaCl or 1 g of KCl?
18. Differentiate between homoatomic and heteroatomic molecules with examples.
19. In which one of the followings the number of hydrogen atoms is more? 2 moles of HCl or 1 mole of NH_3 (Hint:1 mole of a substance contains as much number of moles of atoms as are in 1 molecule of a substance.

Long Answer Questions.

1. Define element and classify the elements with examples.
2. List five characteristics by which compounds can be distinguished from mixtures.
3. Differentiate between the following with examples:
   i. Molecule and gram molecule ii. Atom and gram atom
   iii. Molecular mass and molar mass iv. Chemical formula and gram formula
4. Mole is SI unit for the amount of a substance. Define it with examples?

Numericals

1. Sulphuric acid is the king of chemicals. If you need 5 moles of sulphuric acid for a reaction, how many grams of it will you weigh?
2. Calcium carbonate is insoluble in water. If you have 40 g of it; how many Ca^{2+} and CO_3^{2-} ions are present in it?
3. If you have 6.02×10^{23} ions of aluminium; how many sulphate ions will be required to prepare $Al_2(SO_4)_3$?
4. Calculate the number of molecules in the following compounds:
   a. 16 g of H_2CO_3 b. 20 g of HNO_3 c. 30 g of $C_6H_{12}O_6$
5. Calculate the number of ions in the following compounds:
   a. 10 g of $AlCl_3$ b. 30 g of $BaCl_2$ c. 58 g of $H_2SO_{4(aq)}$
6. What will be the mass of 2.05×10^{16} molecules of H_2SO_4
7. How many atoms are required to prepare 60 g of HNO_3?
8. How many ions of Na^+ and Cl^- will be present in 30 g of NaCl?
9. How many molecules of HCl will be required to have 10 grams of it?
10. How many grams of Mg will have the same number of atoms as 6 grams of C have?

Chapter 2

Structure of Atoms

Major Concepts

2.1 Theories and Experiments related to Atomic Structure
2.2 Electronic Configuration
2.3 Isotopes

Time allocation	
Teaching periods	16
Assessment periods	03
Weightage	10%

Students Learning Outcomes

Students will be able to:

- Describe the contributions that Rutherford made to the development of the Atomic Theory.
- Explain how Bohr's atomic theory differed.
- Describe the structure of atom including the location of the proton, electron and neutron.
- Define isotopes.
- Compare isotopes of an atom.
- Discuss the properties of the isotopes of H, C, Cl, U.
- Draw the structure of different isotopes from mass number and atomic number.
- State the importance and uses of isotopes in various fields of life.
- Describe the presence of subshells in shell.
- Distinguish between shells and subshells.
- Write the electronic configuration of first 18 elements in the Periodic Table.

Introduction

Ancient Greek philosopher Democritus suggested that matter is composed of tiny indivisible particles called atoms. The name atom was derived from the Latin word 'Atomos' meaning indivisible. In the beginning of 19^{th} century John Dalton put forward Atomic Theory. According to it 'all matter is made up of very small indivisible particles called atoms'. Till the end of 19^{th} century it was considered that atom cannot be subdivided. However, in the beginning of 20^{th} century, experiments performed by Goldstein, J. J. Thomson, Rutherford, Bohr and other scientist revealed that atom is made up of subatomic particles like electron, proton and neutron. Properties of these subatomic particles will be discussed in this chapter.

2.1 THEORIES AND EXPERIMENTS RELATED TO STRUCTURE OF ATOM

According to Dalton, an atom is an indivisible, hard, dense sphere. Atoms of the same element are alike. They combine in different ways to form compounds. In the light of Dalton's atomic theory, scientists performed a series of experiments. But in the late 1800's and early 1900's, scientists discovered new subatomic particles.

J.J Thomson (1856-1940) was a British physicist. He was awarded the 1906 Noble Prize in Physics for the discovery of electron and for his work on the conduction of electricity in gases

In 1886, Goldstein discovered positively charged particles called **protons**. In 1897, J.J. Thomson found in an atom, the negatively charged particles known as **electrons**. It was established that electrons and protons are fundamental particles of matter. Based upon these observations Thomson put forth his **"plum pudding"** theory. He postulated that atoms were solid structures of positively charge with tiny negative particles stuck inside. It is like plums in the pudding.

Cathode rays and Discovery of Electron

In 1879, Sir William Crooks performed experiments by passing electric current through gases in a discharge tube at very low pressure.

He took a glass tube fitted with two metallic electrode, which were connected to a high voltage battery. The pressure inside the tube was kept 10^{-4} *atm.* When high voltage current was passed through the gas, shiny rays were emitted from the cathode which travel towards the anode as shown in figure 2.1. These rays were given the name of **'cathode rays'** as these were originated from the cathode.

Sir William Crooks (1832-1919) was a British chemist and physicist. He was pioneer of vacuum tubes. He worked on spectroscopy

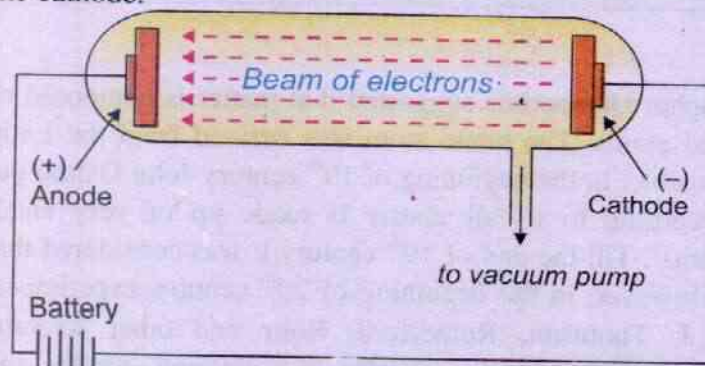



Fig 2.1 Discharge tube used for the production of cathode rays.

The cathode rays were studied in detail and their properties were determined, which are given below:

i. These rays travel in straight lines perpendicular to the cathode surface.
ii. They can cast a sharp shadow of an opaque object if placed in their path.
iii. They are deflected towards positive plate in an electric field showing that they are negatively charged.
iv. They raise temperature of the body on which they fall.
v. J.J. Thomson discovered their charge/mass (e / m) ratio.
vi. Light is produced when these rays hit the walls of the discharge tube.
vii. It was found that the same type of rays were emitted no matter which gas and which cathode was used in the discharge tube.

All these properties suggested that the nature of cathode rays is independent of the nature of the gas present in the discharge tube or material of the cathode. The fact that they cast the shadow of an opaque object suggested that these are not rays but they are fast moving material particles. They were given the name **electrons.** Since all the materials produce same type of particles, it means all the materials contain electrons. As we know materials are composed of atoms, hence the electrons are fundamental particles of atoms.

Discovery of Proton

In 1886, Goldstein observed that in addition to cathode rays, other rays were also present in the discharge tube. These rays were traveling in opposite direction to cathode rays. He used a discharge tube having perforated cathode as shown in figure 2.2. He found that these rays passed through holes present in the cathode and produced a glow on the walls of the discharge tube. He called these rays as "**canal rays**".

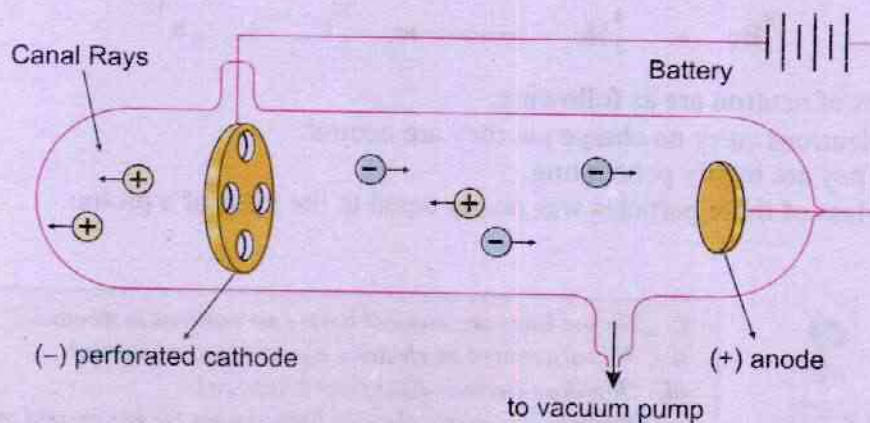



Fig 2.2 Discharge tube used for the production of canal rays.

The properties of these rays were as following:

i. These rays travel in straight lines in a direction opposite to the cathode rays.

ii. Their deflection in electric and magnetic field proved that these are positively charged.

iii. The nature of canal rays depends upon the nature of gas, present in the discharge tube.

iv. These rays do not originate from the anode. In fact these rays are produced when the cathode rays or electrons collide with the residual gas molecules present in the discharge tube and ionize them as follows:

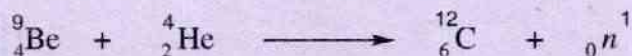
$$M + e^- \longrightarrow M^+ + 2e^-$$

v. Mass of these particles was found equal to that of a proton or simple multiple of it. The mass of a proton is 1840 times more than that of an electron.

Thus, these rays are made up of positively charged particles. The mass and charge of these particles depend upon the nature of the gas in the discharge tube. Hence, different gases produce different types of positive rays having particles of different masses and different charges. Keep in mind that positive particles produced by a gas will be of the same type i.e. positive rays produced by the lightest gas hydrogen contain protons.

Discovery of Neutron

Rutherford observed that atomic mass of the element could not be explained on the basis of the masses of electron and proton only. He predicted in 1920 that some neutral particle having mass equal to that of proton must be present in an atom. Thus scientists were in search of such a neutral particle. Eventually in 1932 Chadwick discovered neutron, when he bombarded alpha particles on a beryllium target. He observed that highly penetrating radiations were produced. These radiations were called neutron.

$${}^{9}_{4}Be + {}^{4}_{2}He \longrightarrow {}^{12}_{6}C + {}_{0}n^{1}$$

Properties of neutron are as following:

i. Neutrons carry no charge i.e. they are neutral.

ii. They are highly penetrating.

iii. Mass of these particles was nearly equal to the mass of a proton.

Test yourself 2.1

i. Do you know any element having no neutrons in its atoms?

ii. Who discovered an electron, a proton and a neutron?

iii. How does electron differ from a neutron?

iv. Explain, how anode rays are formed from the gas present in the discharge tube?

2.1.1 Rutherford's Atomic Model

Rutherford performed 'Gold Foil' experiment to understand how negative and positive charges could coexist in an atom. He bombarded *alpha* particles on a 0.00004 *cm* thick gold foil. Alpha particles are emitted by radioactive elements like radium and polonium. These are actually helium nuclei (He^{2+}). They can penetrate through matter to some extent.

He observed the effects of α-particles on a photographic plate or a screen coated with zinc sulphide as shown in figure 2.3. He proved that the 'plum-pudding' model of the atom was not correct.

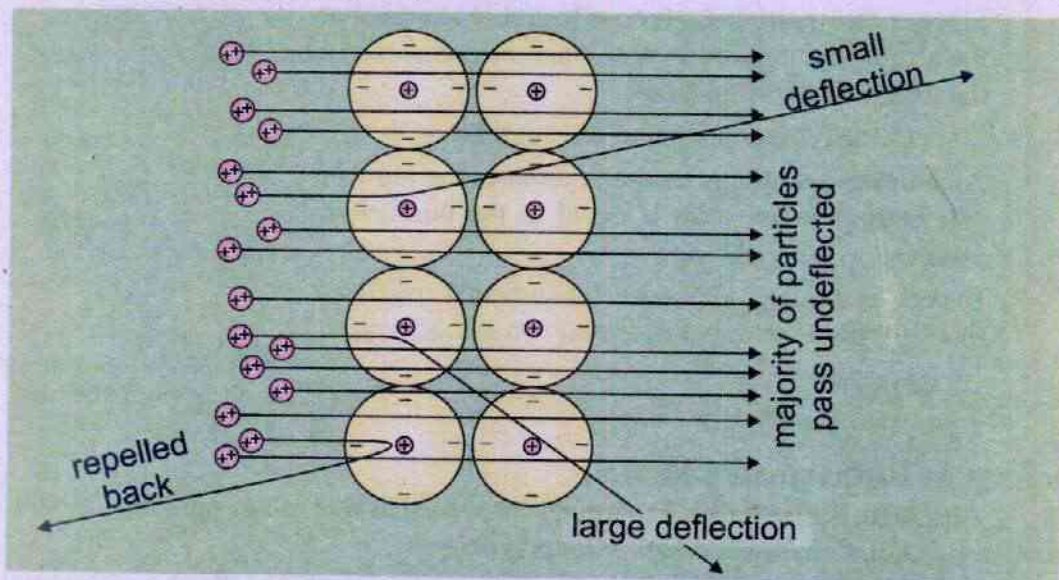



Fig 2.3 Scattering of alpha particles by the atoms of gold foil.

Observations made by Rutherford were as follows:

i. Almost all the particles passed through the foil undeflected.

ii. Out of 20000 particles, only a few were deflected at fairly large angles and very few bounced back on hitting the gold foil.

Results of the experiment

Keeping in view the experiment, Rutherford proposed planetary model for an atom and concluded following results:

i. Since most of the particles passed through the foil undeflected, therefore most of the volume occupied by an atom is empty .

ii. The deflection of a few particles proved that there is a 'center of positive charges' in an atom, which is called 'nucleus' of an atom.

iii. The complete rebounce of a few particles show that the nucleus is very dense and hard.

iv. Since a few particles were deflected it shows that the size of the nucleus is very small as compared to the total volume of an atom.

v. The electrons revolve around the nucleus.

vi. An atom as a whole is neutral, therefore the number of electrons in an atom is equal to the number of protons.

vii. Except electrons, all other fundamental particles that lie within the nucleus, are known as nucleons.

Rutherford was a British-New Zealand chemist. He performed a series of experiments using α – particles. He won the 1908 Noble Prize in Chemistry. In 1911, he proposed the nuclear model of the atom and performed the first experiment to split atom. Because of his great contributions, he is considered the father of nuclear science.

Defects in Rutherford's Model

Although Rutherford's experiment proved that the 'plum-pudding' model of an atom was not correct, yet it had following defects:

i. According to classical theory of radiation, electrons being the charged particles should release or emit energy continuously and they should ultimately fall into the nucleus.

ii. If the electrons emit energy continuously, they should form a continuous spectrum but in fact, line spectrum was observed.

Although the scientists had objections on the atomic model presented by Rutherford, yet it cultivated thought provoking ideas among them. They initiated the quest to answer the following questions:

i. How can an atom collapse or why are atoms stable?

ii. Why does an atom give line spectrum?

iii. Scientists considered there must be another model of atom.

It indicated that Rutherford's model was not perfect.

2.1.2 Bohr's Atomic Theory

Keeping in view the defects in Rutherford's Atomic Model, Niels Bohr presented another model of atom in 1913. The Quantum Theory of Max Planck was used as foundation for this model. According to Bohr's model revolving electron in an atom does not absorb or emit energy continuously. The energy of a revolving electron is 'quantized' as it revolves only in orbits of fixed energy, called 'energy levels' by him. The Bohr's atomic model is shown in figure 2.4.

Niels Bohr was a Danish physicist who joined Rutherford in 1912 for his post doctoral research. In 1913, Bohr presented his atomic model based upon Quantum theory. He won the 1922 Noble Prize for Physics for his work on the structure of an atom.

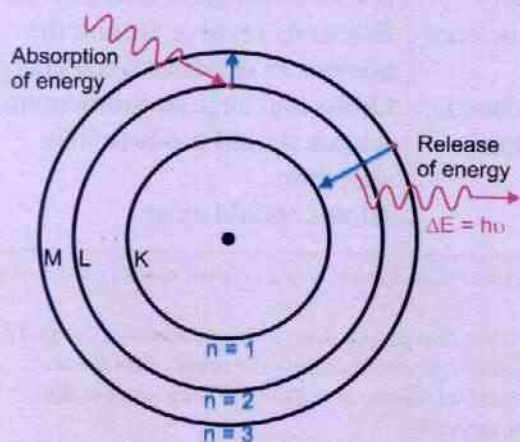



Fig 2.4 Bohr's atomic model showing orbits.

The Bohr's atomic model was based upon the following postulates:

i. The hydrogen atom consists of a tiny nucleus and electrons are revolving in one of circular orbits of radius '*r*' around the nucleus.

ii. Each orbit has a fixed energy that is quantized.

iii. As long as electron remains in a particular orbit, it does not radiate or absorb energy. The energy is emitted or absorbed only when an electron jumps from one orbit to another.

iv. When an electron jumps from lower orbit to higher orbit, it absorbs energy and when it jumps from higher orbit to lower orbit it radiates energy. This change in energy, $\triangle E$ is given by following Planck's equation

$$\triangle E = E_2 - E_1 = h\nu$$

Where, h is Planck's constant equal to 6.63×10^{-34} Js, and ν is frequency of light.

v. Electron can revolve only in orbits of a fixed angular moment mvr, given as:

$$mvr = n\frac{h}{2\pi}$$

Where 'n' is the quantum number or orbit number having values 1,2,3 and so on.

Do you know?

Quantum means fixed energy. It is the smallest amount of energy that can be emitted or absorbed as electromagnetic radiation. Quanta is plural of quantum.

In 1918 Noble prize in physics was awarded to German physicist Max Planck (1858-1947) for his work on the quantum theory

Summary of differences between two theories:

	Rutherford's Atomic Theory	Bohr's Atomic Theory
i.	It was based upon classical theory.	It was based upon quantum theory.
ii.	Electrons revolve around the nucleus.	Electrons revolve around the nucleus in orbits of fixed energy.
iii.	No idea about orbits was introduced.	Orbits had angular momentum.
iv.	Atoms should produce continuous spectrum.	Atoms should produce line spectrum.
v.	Atoms should collapse.	Atoms should exist.

Test yourself 2.2

i. How was it proved that the whole mass of an atom is located at its centre?
ii. How was it shown that atomic nuclei are positively charged?
iii. Name the particles which determine the mass of an atom.
iv. What is the classical theory of radiation? How does it differ from quantum theory?
v. How can you prove that angular momentum is quantized?

Hint: *Let angular momentum (mvr) of 1st orbit is $mvr = nh/2\pi$*
By putting the values of h and π

$$mvr = \frac{6.63 \times 10^{-34}}{2 \times 3.14} = 1.0 \times 10^{-34}\ kg\ m^2\ s^{-1}$$

2.2 ELECTRONIC CONFIGURATION

Before discussing electronic configuration let us first understand the concept of shells and subshells.

We have learnt about the structure of atom i.e. it consists of a tiny nucleus lying in the center and electrons revolving around the nucleus. Now we will discuss how the electrons revolve around the nucleus? The electrons revolve around the nucleus in different energy levels or shells according to their respective energies (potential energy). The concept of potential energy of an electron shall be discussed in higher classes.

Energy levels are represented by 'n' values 1, 2, 3 and so on. They are designated by the alphabets K, L, M and so on. A shell closer to the nucleus is of minimum energy. Since K shell is closest to the nucleus, the energy of shells increases from K shell onwards. Such as:

1st energy level is K shell; it has the lowest energy.
2nd energy level is L shell; it has more energy than K shell.
3rd energy level is M shell; it has more energy than K and L shell.
4th energy level is N shell; it has more energy than K, L and M shell.

In simple words shells are the main energy levels that electrons can occupy. Shells are represented by circles around the nucleus. They are counted from the center to outward as shown in the figure 2.5. The number of electrons that a shell can accommodate is given by $2n^2$, where 'n' is the shell number.

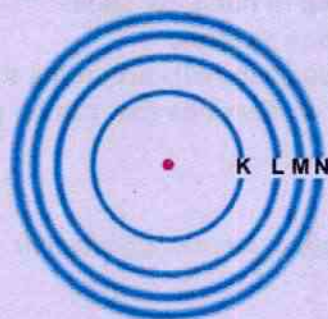



Fig. 2.5 Showing different energy levels or shells counted from the centre outwards.

A shell also consists of subshells the number of subshells in a shell is equal to its n value. Each subshell is designated by a small alphabetical letter ***s, p, d, f*** etc.

First energy level or K shell has only one subshell called s subshell. Second energy level or L shell has two subshells *s* and *p*. Third energy level M shell has three subshells *s*, *p* and *d*. Fourth energy level or N shell has four subshells *s*, *p*, *d* and *f*.

n value	Shell	Subshell
1	K	Only s
2	L	*s, p*
3	M	*s, p, d*
4	N	*s, p, d, f*

2.2.1 Electronic Configuration of First 18 Elements

The distribution of electrons around the nucleus in various shells and subshells according to their increasing energy, is called **electronic configuration**. The most stable or ground state electronic configuration of an atom is the one in which electrons are present in the lowest possible energy level. The electrons fill the shells in order of their increasing energy, i.e. lower energy level is occupied first then the higher energy level and so on as indicated earlier. The maximum capacity of shells to accommodate the electrons is as following:

Fig. 2.8 Filling sequence of electrons in subshells according to energy levels

K shell can accommodate 2 electrons.
L shell can accommodate 8 electrons.
M shell can accommodate 18 electrons.

N shell can accommodate 32 electrons.

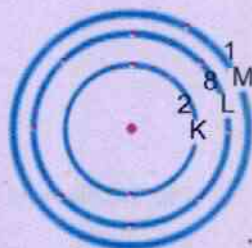
As we know there is a slight difference between the energies of the subshells within a shell, therefore, filling of electrons in subshells of a shell is such as that '*s*' subshell is filled first and then its *p* subshell and then other subshells are filled. The maximum capacity of subshells to accommodate electrons is:

'*s*' subshell can accommodate 2 electrons.
'*p*' subshell can accommodate 6 electrons.

Let us write the electronic configuration of the elements and their ions with the help of a few examples. Keep in mind, we should know three things:

i. The number of electrons in an atom.
ii. The sequence of shells and subshells according to the energy levels.
iii. The maximum number of electrons that can be placed in different shells and subshells.

Example 2.1

Write the electronic configuration of an element having 11 electrons.

Solution:

Keep in mind that all electrons do not have the same energy. Therefore, they are accommodated in different shells according to increasing energy and capacity of the shell. First of all, the electrons will go to K shell which has minimum energy. It can accommodate 2 electrons. After this, electrons will go to L shell that can accommodate 8 electrons. Thus K and L shells have accommodated 10 electrons. The remaining 1 electron will go to M shell, the outermost shell of maximum energy in this case.

The electronic configuration will be written as: K L M

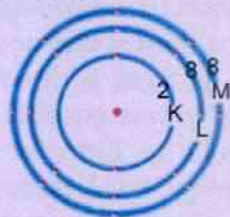
2, 8, 1,

But it is not necessary to write the subshells. Therefore, it is simply written as 2,8,and 1. Further distribution of electrons in subshells will be: $1s^2$, $2s^2$, $2p^6$, $3s^1$.

Example 2.2

Write down the electronic configuration of Cl^- ion

Solution:

We know that chlorine has 17 electrons and chloride ion (Cl^-) has 17 + 1 = 18 electrons. Its electronic configuration will be 2, 8, 8, which is presented in the figure. The further distribution of electrons in subshells will be $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$.

Example 2.3

An element has 5 electrons in M shell. Find out its atomic number.

Solution:

When there are 5 electrons in M shell, it means K and L shell are completely filled with their maximum capacity of 10 electrons. Hence the electronic configuration of the element is: K L M

2, 8, 5, or just 2, 8, 5

So, the total number of electrons is $2 + 8 + 5 = 15$

As we know, the number of electrons in an atom is equal to its atomic number. Therefore, atomic number of this element is 15.

2.2.2 The electronic configuration of first 18 elements

The sequence of filling of electrons in different subshells is as following:

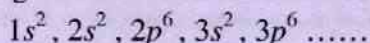
$$1s^2, 2s^2, 2p^6, 3s^2, 3p^6 \ldots\ldots$$

Where number represents the shell number, while letters (*s* and *p*) represent subshells. The superscript shows the number of electrons in a subshell. The sum of superscripts number is the total number of electrons in an atom. i.e. atomic number of an element. The electronic configuration of first 18 elements is shown in the Table 2.1

Table 2.1 Electronic Configuration of First Eighteen Elements

Element	Symbol	Atomic Number	Electronic Configuration
Hydrogen	H	1	$1s^1$
Helium	He	2	$1s^2$
Lithium	Li	3	$1s^2, 2s^1$
Beryllium	Be	4	$1s^2, 2s^2$
Boron	B	5	$1s^2, 2s^2, 2p^1$
Carbon	C	6	$1s^2, 2s^2, 2p^2$
Nitrogen	N	7	$1s^2, 2s^2, 2p^3$
Oxygen	O	8	$1s^2, 2s^2, 2p^4$
Fluorine	F	9	$1s^2, 2s^2, 2p^5$
Neon	Ne	10	$1s^2, 2s^2, 2p^6$
Sodium	Na	11	$1s^2, 2s^2, 2p^6, 3s^1$
Magnesium	Mg	12	$1s^2, 2s^2, 2p^6, 3s^2$
Aluminium	Al	13	$1s^2, 2s^2, 2p^6, 3s^2, 3p^1$
Silicon	Si	14	$1s^2, 2s^2, 2p^6, 3s^2, 3p^2$
Phosphorus	P	15	$1s^2, 2s^2, 2p^6, 3s^2, 3p^3$
Sulphur	S	16	$1s^2, 2s^2, 2p^6, 3s^2, 3p^4$
Chlorine	Cl	17	$1s^2, 2s^2, 2p^6, 3s^2, 3p^5$
Argon	Ar	18	$1s^2, 2s^2, 2p^6, 3s^2, 3p^6$

Test yourself 2.3

i. *How many the maximum number of electrons that can be accommodated in a p-subshell?*
ii. *How many subshells are there in second shell?*
iii. *Why does an electron first fill 2p orbital and then 3s orbital?*
iv. *If both K and L shells of an atom are completely filled; how many total number of electrons are present in them?*
v. *How many electrons can be accommodated in M shell?*
vi. *What is the electronic configuration of a hydrogen atom?*
vii. *What is atomic number of phosphorus? Write down its electronic configuration.*
viii. *If an element has atomic number 13 and atomic mass 27; how many electrons are there in each atom of the element?*
ix. *How many electrons will be in M shell of an atom having atomic number 15,*
x. *What is maximum capacity of a shell?*

2.3 ISOTOPES

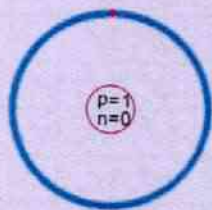
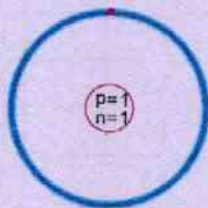
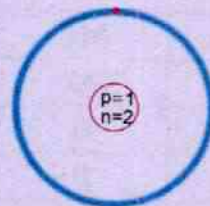
2.3.1 Definition

Isotopes *are defined as the atoms of an element that have same atomic number but different mass numbers*. They have same electronic configuration and number of protons but they differ in the number of neutrons. Isotopes have similar chemical properties because these depend upon electronic configuration. But they have different physical properties because these depend upon mass numbers. Most of the elements have isotopes. Here we will discuss the isotopes of hydrogen, carbon, chlorine and uranium only.

2.3.2 Examples

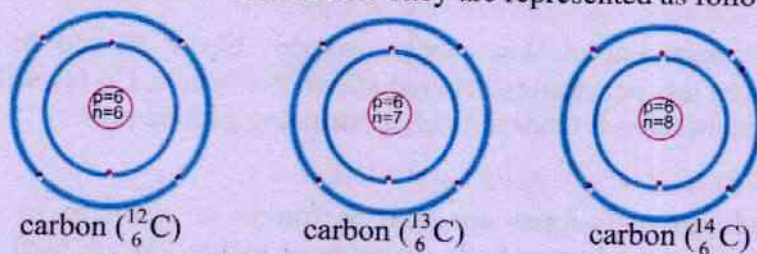
i) Isotopes of Hydrogen

The naturally occurring hydrogen is combination of its three isotopes, present in different abundances. The three isotopes of hydrogen are named as **protium** (^1_1H), **deuterium** (^2_1H, or D) and **tritium** (^3_1H or T). Each one of them has 1 proton and 1 electron, but number of neutrons are different as shown in Table 2.2

The isotopes are represented as:

protium (^1_1H) deuterium (^2_1H) tritium (^3_1H)

ii) Isotopes of Carbon

There are two stable isotopes of carbon ^{12}C and ^{13}C and one radioactive isotope ^{14}C. The isotope ^{12}C is present in abundance of 98.9 %, while ^{13}C and ^{14}C are both present only 1.1 % in nature. All of them have the same number of protons and electrons but differ in number of neutrons. They are represented as follows:

carbon ($^{12}_{6}C$) carbon ($^{13}_{6}C$) carbon ($^{14}_{6}C$)

iii) Isotopes of Chlorine

There are two isotopes of chlorine, $^{35}_{17}Cl$ and $^{37}_{17}Cl$

iv) Isotopes of Uranium

There are three isotopes of uranium i.e. $^{234}_{92}U$, $^{235}_{92}U$ and $^{238}_{92}U$. The $^{238}_{92}U$ is found in nature nearly 99%.

The difference in their number of electrons, protons and neutrons is shown below:

Table 2.2 Atomic Number, Mass Number, Number of Protons and Neutrons of H, C, Cl and U

Symbol	Atomic Number	Mass Number	No. of Proton	No. of Neutron
^{1}H	1	1	1	0
^{2}H	1	2	1	1
^{3}H	1	3	1	2
^{12}C	6	12	6	6
^{13}C	6	13	6	7
^{14}C	6	14	6	8
^{35}Cl	17	35	17	18
^{37}Cl	17	37	17	20
^{234}U	92	234	92	142
^{235}U	92	235	92	143
^{238}U	92	238	92	146

APPLICATION OF ISOTOPES

In science and many different technological fields isotopes have vast applications. The biggest application is in the field of medicine. They are applied in diagnosis, radiotherapy and treatment of many diseases like cancer.

2.3.3 Uses

With the advancement of the scientific knowledge, the isotopes have found many applications in our lives. Following are the major fields in which isotopes have vast applications:

i. Radiotherapy (Treatment of Cancer)

For the treatment of skin cancer, isotopes like P-32 and Sr-90 are used because they emit less penetrating beta radiations. For cancer, Co-60, affecting within the body, is used because it emits strongly penetrating gamma rays.

ii. Tracer for Diagnosis and Medicine

The radioactive isotopes are used as tracers in medicine to diagnose the presence of tumor in the human body. Isotopes of Iodine-131 are used for diagnosis of goiter in thyroid gland. Similarly technetium is used to monitor the bone growth.

iii. Archaeological and Geological Uses

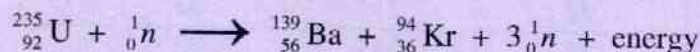
The radioactive isotopes are used to estimate the age of fossils like dead plants and animals and stones etc. The age determination of very old objects based on the half-lives of the radioactive isotope is called radioactive-isotope dating. An important method of *age determination of old carbon containing objects (fossils) by measuring the radioactivity of C-14 in them is called radio-carbon dating or simply* ***carbon dating***.

iv. Chemical Reaction and Structure Determination

The radioisotopes are used in a chemical reaction to follow a radioactive element during the reaction and ultimately to determine the structure. For example: C-14 is used to label CO_2. As CO_2 is used by the plants for photosynthesis to form glucose, its movement is detected through the various intermediate steps up to glucose.

v. Applications in Power Generation

Radioactive isotopes are used to generate electricity by carrying out controlled nuclear fission reactions in nuclear reactors. For example, when U-235 is bombarded with slow moving neutrons, the uranium nucleus breaks up to produce Barium-139 and Krypton-94 and three neutrons.

$$^{235}_{92}\text{U} + {}^{1}_{0}n \longrightarrow {}^{139}_{56}\text{Ba} + {}^{94}_{36}\text{Kr} + 3\,{}^{1}_{0}n + \text{energy}$$

A large amount of energy is released which is used to convert water into steam in boilers. The steam then drives the turbines to generate electricity. This is the peaceful use of atomic energy for development of a nation.

Test yourself 2.4

i. *Why do the isotopes of an element have different atomic masses?*
ii. *How many neutrons are present in C-12 and C-13?*
iii. *Which of the isotopes of hydrogen contains greater number of neutrons?*
iv. *Give one example each of the use of radioactive isotope in medicine and radiotherapy.*
v. *How is the goiter in thyroid gland detected?*
vi. *Define nuclear fission reaction.*
vii. *When U-235 breaks up, it produces a large amount of energy. How is this energy used?*
viii. *How many neutrons are produced in the fission reaction of U-235?*
ix. *U-235 fission produces two atoms of which elements?*

TESTING PREVAILING THEORIES BRINGS ABOUT CHANGE IN THEM

Science is a process for producing knowledge. The process depends both on making careful observations of phenomenae and inventing theories for making sense out of those observations. Change in knowledge is inevitable because new observations may challenge prevailing theories. No matter how well one theory explains a set of observations, it is possible that another theory may fit just as well or better, or may fit a still wider range of observations. In science, the testing and improving and occasional discarding of theories, whether new or old, go on all the time. Scientists assume that even if there is no way to secure complete and absolute truth, increasingly accurate approximations can be made to account for the world and how it works.

Key Points

- Cathode rays were discovered in last decade of nineteen century. The properties of cathode rays were determined and they led to the discovery of electron.
- Canal rays were discovered in 1886 by Goldstein . The properties of canal rays resulted in the discovery of proton in the atom.
- Neutron in the atom was discovered in 1932 by Chadwick.
- First of all structure of an atom was presented by Rutherford in 1911. He proposed that an atom contains nucleus at the center and electrons revolve around this nucleus.
- Bohr presented an improved model of an atom in 1913 based upon four postulates. He introduced the concept of circular orbit, in which electrons revolve. As long as electron remains in a particular orbit, it does not radiate energy. Release and gain of energy is because of change of orbit.
- The concept of shells and subshells is explained.
- A ***shell*** consists of ***subshells***.
- Isotopes are defined as the atoms of elements that have the same atomic number but different atomic mass.
- Hydrogen, carbon and uranium have three isotopes each, whereas chlorine has two isotopes.

EXERCISE

Multiple Choice Questions

Put a (✓) on the correct answer

1. **Which one of the following results in the discovery of proton**
   (a) cathode rays (b) canal rays (c) X-rays (d) alpha rays.
2. **Which one of the following is the most penetrating.**
   (a) protons (b) electrons (c) neutrons (d) alpha particles
3. **The concept of orbit was used by**
   (a) J. J. Thomson (b) Rutherford (c) Bohr (d) Planck
4. **Which one of the following shell consists of three subshells.**
   (a) O shell (b) N shell (c) L shell (d) M shell
5. **Which radioisotope is used for the diagnosis of tumor in the body?**
   (a) cobalt-60 (b) iodine-131 (c) strontium-90 (d) phosphorus- 32
6. **When U-235 breaks up, it produces:**
   (a) electrons (b) neutrons (c) protons (d) nothing
7. **The p subshell has:**
   (a) one orbital (b) two orbitals (c) three orbitals (d) four orbitals
8. **Deuterium is used to make:**
   (a) light water (b) heavy water (c) soft water (d) hard water
9. **The isotope C-12 is present in abundance of:**
   (a) 96.9 % (b) 97.6 % (c) 99.7 % (d) none of these
10. **Who discovered the proton:**
    (a) Goldstein (b) J. J. Thomson (c) Niels Bohr (d) Rutherford

Short Answer Questions

1. What is the nature of charge on cathode rays?
2. Give five characteristics of cathode rays.
3. The atomic symbol of a phosphorus ion is given as $^{31}_{15}P^{3-}$
   (a) How many protons, electrons and neutrons are there in the ion?
   (b) What is name of the ion?
   (c) Draw the electronic configuration of the ion.
   (d) Name the noble gas which has the same electronic configuration as the phosphorus ion has.
4. Differentiate between shell and subshell with examples of each.
5. An element has an atomic number 17. How many electrons are present in K, L and M shells of the atom?

6. Write down the electronic configuration of Al^{3+}. How many electrons are present in its outermost shell?
7. Magnesium has electronic configuration 2, 8, 2,
   (a) How many electrons are in the outermost shell?
   (b) In which subshell of the outermost shell electrons are present?
   (c) Why magnesium tends to lose electrons?
8. What will be the nature of charge on an atom when it loses an electron or when it gains an electron?
9. For what purpose U-235 is used?
10. A patient has goiter. How will it be detected?
11. Give three properties of positive rays.
12. What are the defects of Rutherford's atomic model?
13. As long as electron remains in an orbit, it does not emit or absorb energy. When does it emit or absorb energy?

Long Answer Questions

1. How are cathode rays produced? What are its five major characteristics?
2. How was it proved that electrons are fundamental particles of an atom?
3. Draw a labeled diagram to show the presence of protons in the discharge tube and explain how canal rays were produced.
4. How Rutherford discovered that atom has a nucleus located at the centre of the atom?
5. One of the postulates of Bohr's atomic model is that angular momentum of a moving electron is quantized. Explain its meaning and calculate the angular momentum of third orbit (i.e. n=3)
6. How did Bohr prove that an atom must exist?
7. What do you mean by electronic configuration? What are basic requirements while writing electronic configuration of an element (atom)?
8. Describe the electronic configuration of Na^{+}, Mg^{2+} and Al^{3+} ions. Do they have the same number of electrons in the outermost shell?
9. Give the applications of isotopes in the field of radiotherapy and medicines.
10. What is an isotope? Describe the isotopes of hydrogen with diagrams.

Periodic Table and Periodicity of Properties

Major Concepts

3.1 Periodic Table
3.2 Periodic Properties

Time allocation	
Teaching periods	12
Assessment periods	02
Weightage	10%

Students Learning Outcomes

Students will be able to:

- Distinguish between period and group in the Periodic table.
- State the Periodic law.
- Classify elements (into two categories: groups and periods) according to the configuration of their outermost electrons.
- Determine the demarcation of the periodic table into *s*-block and *p*-block.
- Explain the shape of the periodic table.
- Determine the location of families of the periodic table.
- Recognize the similarity in the physical and chemical properties of elements in the same family of the elements.
- Identify the relationship between electronic configuration and position of elements in the periodic table.
- Explain how shielding effect influences periodic trends.
- Describe how electronegativities change within a group and within a period in the periodic table.

Introduction

In nineteenth century chemists devoted much of their efforts in attempts to arrange elements in a systematic manner. These efforts resulted in discovery of periodic law. On the basis of this law, the elements known at that time, were arranged in the form of a table which is known as periodic table. One of the significant features of the table was that it predicted the properties of those elements which were not even discovered at that time. The vertical columns of that table were called groups and

horizontal lines were called periods. That orderly arrangement of elements generally coincided with their increasing atomic number. The periodic table contains huge amount of information for scientists.

3.1 PERIODIC TABLE

With the discovery of the periodic table the study of individual properties of the known elements is reduced to study of a few groups. We will describe various attempts which were made to classify the elements into a tabular form.

Dobereiner's Triads

A German chemist Dobereiner observed relationship between atomic masses of several groups of three elements called ***triads.*** In these groups, the central or middle element had atomic mass average of the other two elements. One triad group example is that of calcium (40), strontium(88) and barium (137). The atomic mass of strontium is the average of the atomic masses of calcium and barium. Only a few elements could be arranged in this way. This classification did not get wide acceptance.

Newlands Octaves

After successful determination of correct atomic masses of elements by Cannizzaro in 1860, attempts were again initiated to organize elements. In 1864 British chemist Newlands put forward his observations in the form of '*law of octaves*'. He noted that there was a repetition in chemical properties of every eighth element if they were arranged by their increasing atomic masses. He compared it with musical notes. His work could not get much recognition as no space was left for undiscovered element. The noble gases were also not known at that time.

Mendeleev (1834-1907) was a Russian chemist and inventor. He was the creator of first version of periodic table of elements. With help of the table, he predicted the properties of elements yet to be discovered.

Mendeleev's Periodic Table

Russian chemist, Mendeleev arranged the known elements (only 63) in order of increasing atomic masses, in horizontal rows called periods. So that elements with similar properties were in the same vertical columns. This arrangement of elements was called Periodic Table. He put forward the results of his work in the form of periodic law, which is stated as "*properties of the elements are periodic functions of their atomic masses*"

Although, Mendeleev periodic table was the first ever attempt to arrange the elements, yet it has a few demerits in it. His failure to explain the position of isotopes and wrong order of the atomic masses of some elements suggested that atomic mass of an element cannot serve as the basis for the arrangement of elements.

Periodic Law

In 1913 H. Moseley discovered a new property of the elements i.e. atomic number. He observed that atomic number instead of atomic mass should determine the position of element in the periodic table and accordingly the **periodic law** was amended as *"properties of the elements are periodic function of their atomic numbers"*. Atomic number of an element is equal to the number of electrons in a neutral atom. So atomic number provides the basis of electronic configurations as well.

Do you know?

Atomic number is a more fundamental property than atomic mass because atomic number of every element is fixed and it increases regularly by 1 from element to element. No two elements can have the same atomic number.

Test yourself 3.1

i. What was the contribution of Dobereiner towards classification of elements?
ii. How Newlands arranged the elements?
iii. Who introduced the name Periodic Table?
iv. Why the improvement in Mendeleev's periodic table was made?
v. State Mendeleev's periodic law.
vi. Why and how elements are arranged in a period?

Modern Periodic Table

Atomic number of an element is more fundamental property than atomic mass in two respects. (a) It increases regularly from element to element, (b) It is fixed for every element. So the discovery of atomic number of an element in 1913 led to change in Mendeleev's periodic law which was based on atomic mass.

The modern periodic table is based upon the arrangement of elements according to increasing atomic number. When the elements are arranged according to increasing atomic number from left to right in a horizontal row, *properties of elements were found repeating after regular intervals* such that elements of similar properties and similar configuration are placed in the same group.

It was observed that after every eighth element, ninth element had similar properties to the first element. For example, sodium (Z=11) had similar properties to lithium (Z=3). After atomic number 18, every nineteenth element was showing similar behaviour. So the long rows of elements were cut into rows of eight and eighteen elements, and placed one above the other so that a table of vertical and horizontal rows was obtained.

Long form of Periodic Table

The significance of atomic number in the arrangement of elements in the modern periodic table lies in the fact that as electronic configuration is based upon atomic number, so the arrangement of elements according to increasing atomic number shows the periodicity (repetition of properties after regular intervals) in the electronic configuration of the elements that leads to periodicity in their properties. Hence the arrangement of elements based on their electronic configuration created a long form of periodic table as shown in figure 3.1.

The horizontal rows of elements in the periodic table are called ***periods***. The elements in a period have continuously increasing atomic number i.e. continuously changing electronic configuration along a period. As a result properties of elements in a period are continuously changing. The number of valence electrons decides the position of an element in a period. For example, elements which have 1 electron in their valence shell occupies the left most position in the respective periods, such as alkali metals. Similarly the elements having 8 electrons in their valence shells such as noble gases always occupy the right most position in the respective periods.

The vertical columns in the periodic table are called ***groups***. These groups are numbered from left to right as 1 to 18. The elements in a group do not have continuously increasing atomic numbers. Rather the atomic numbers of elements in a group increase with irregular gaps.

But the elements of a group have similar electronic configuration i.e. same number of electrons are present in their valence shells. For example, the first group elements have only l electron in their valence shells. Similarly group 2 elements have 2 electrons in their valence shells. It is the reason due to which elements of a group have similar chemical properties.

Salient Features of Long Form of Periodic Table:

i. This table consists of seven horizontal rows called periods

ii. First period consists of only two elements. Second and third periods consist of 8 elements each. Fourth and fifth periods consist of 18 elements each. Sixth period has 32 elements while seventh period has 23 elements and is incomplete, since new elements are expected to be disconcerted.

iii. Elements of a period show different properties.

iv. There are 18 vertical columns in the periodic table numbered 1 to 18 from left to right, which are called groups.

v. The elements of a group show similar chemical properties.

vi. Elements are classified into four blocks depending upon the type of the subshell which gets the last electron.

Fig. 3.1 Modern Periodic Table or long form of the Periodic Table of Elements.

On the basis of completion of a particular subshell, elements with similar subshell electronic configuration are referred as a ***block*** *of elements.* There are four blocks in the periodic table named after the name of the subshell which is in the process of completion by the electrons. These are *s*, *p*, *d* and *f* blocks as shown in figure 3.2. For example, elements of group1 and 2 have valence electrons in '*s*' subshell. Therefore, they are called *s*-block elements as shown in figure 3.2.

Elements of group 13 to 18 have their valence electrons in '*p*' subshell. Therefore, they are referred as p-block elements. The *d*-block lies between the *s* and *p* blocks. While *f*-block lies separately at the bottom. *d*-block constitutes period 4,5 and 6. Each period consists of ten groups starting from group 3 to group 12. These are called **transition metals**.

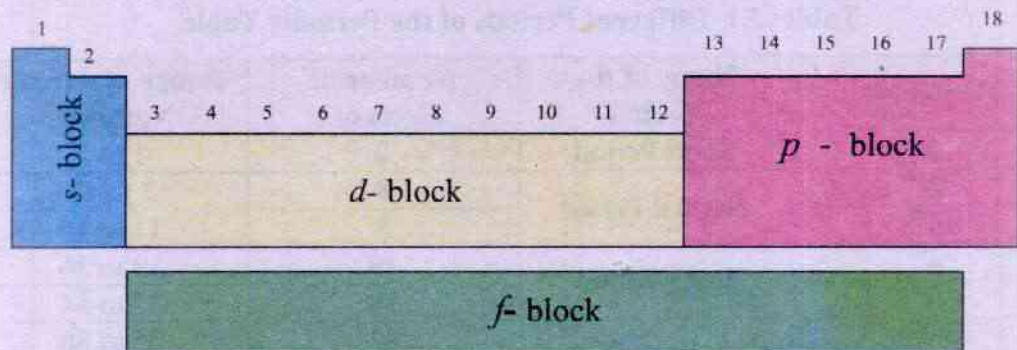



Fig. 3.2 Modern Periodic Table showing four blocks.

Do you know?

Alchemy! *For thousand years alchemy remained field of interest for the scientists. They worked with two main objectives; change common metals into gold and second find cure to diseases and give eternal life to people. They believed all kinds of matter were same combination of four basic elements. Substances are different because these elements combine differently. Changing composition or ratio of any one element, new substances can be formed. The way of making gold from silver or lead was never found and secret of eternal life was never discovered. However, many methods and techniques invented by alchemists are still used in chemistry.*

3.1.1 Periods

First period is called short period. It consists of only two elements, hydrogen and helium. Second and third periods are called normal periods. Each of them has eight elements in it. Second period consists of lithium, beryllium, boron, carbon, nitrogen, oxygen, fluorine and ends at neon, a noble gas. Fourth and fifth periods are called long periods. Each one of them consists of eighteen elements.

Whereas, sixth and seventh periods are called very long periods. In these periods after atomic number 57 and 89, two series of fourteen elements each, were accommodated. Because of space problem, these two series were placed separately below the normal periodic table to keep it in a manageable and presentable form. Since the two series start after Lanthanum (Z=57) and Actinium (Z=89), so these two series of elements are named as Lanthanides and Actinides respectively. Table 3.1 shows the distribution of elements in periods.

All the periods except the first period start with an alkali metal and end at a noble gas. It is to be observed that number of elements in a period is fixed because of maximum number of electrons which can be accommodated in the particular valence shell of the elements.

Table 3.1 Different Periods of the Periodic Table

Period No.	Name of the Period	Number of Elements	Range of Atomic Numbers
1st	Short Period	2	1 to 2
2nd	Normal Period	8	3 to 10
3rd		8	11 to 18
4th	Long Period	18	19 to 36
5th		18	37 to 54
6th	Very Long Period	32	55 to 86
7th		[32]*	87 to 118*

**Since new elements are expected to be discovered, it is an incomplete period*

3.1.2 Groups

Group 1 consists of hydrogen, lithium, sodium, potassium, rubidium, cesium and francium. Although elements of a group do not have continuously increasing atomic numbers, yet they have similar electronic configuration in their valence shells. That is the reason elements of a group are also called a family. For example, all the group 1 elements have one electron in their valence shells, they are given the family name of alkali metals.

The groups 1 and 2 and 13 to 17 contain the normal elements. In the normal elements, all the inner shells are completely filled with electrons, only the outermost shells are incomplete. For example, group 17 elements (halogens) have 7 electrons in their valence (outermost) shell.

The groups 3 to 12 are called transition elements. In these elements '*d*' sub-shell is in the process of completion. Table 3.2 shows the distribution of elements in groups.

Table 3.2 Different Groups of the Periodic Table

Valence electrons	Group number	Family name	General Electronic configuration
1 electron	1	Alkali metals	ns^1
2 electrons	2	Alkaline earth metals	ns^2
3 electrons	13	Boron family	$ns^2\ np^1$
4 electrons	14	Carbon family	$ns^2\ np^2$
5 electrons	15	Nitrogen family	$ns^2\ np^3$
6 electrons	16	Oxygen family	$ns^2\ np^4$
7 electrons	17	Halogen family	$ns^2\ np^5$
8 electrons	18	Noble gases	$ns^2\ np^6$

Do you know?

Fire Works

Beautiful fireworks display are common on celebrations like Pakistan Day or even on marriages. A technology invented in China is used a l l over the world. It is dangerous but careful use of various elements and particularly metal salts of different composition give beauty and colors to the fireworks. Elements like magnesium, aluminium are used in powdered form. Salts of sodium give yellow color, calcium - red; strontium-scarlet; barium-green and copper-bluish green. Usually nitrates and chlorates are used. Other chemicals are added to give brilliance and different shades. Because of fire hazard and risk to life and property, only skilled professionals use them.

Test yourself 3.3

- i. *How the properties of elements repeat after regular intervals?*
- ii. *In which pattern modern periodic table was arranged?*
- iii. *How many elements are in first period and what are their names and symbols?*
- iv. *How many elements are placed in 4^{th} period?*
- v. *From which element lanthanide series starts?*
- vi. *From which period actinides series starts?*
- vii. *How many elements are in 3^{rd} period, write their names and symbols?*
- viii. *How many periods are considered normal periods?*
- ix. *What do you mean by a group in a periodic table?*
- x. *What is the reason of arranging elements in a group?*
- xi. *What do you mean by periodic function?*
- xii. *Why the elements are called s or p block elements?*
- xiii. *Write down the names of elements of group 1 with their symbols?*
- xiv. *How many members are in group 17, is there any liquid, what is its name?*

3.2 PERIODICITY OF PROPERTIES

3.2.1 Atomic Size and Atomic Radius

As we know that atoms are very small and don't have defined boundaries that fix their size. So it is difficult to measure the size of an atom. Therefore, the common method to determine the size of an atom is to assume that atoms are spheres. When they lie close to each other, they touch each other.

Half of the distance between the nuclei of the two bonded atoms is referred as the ***atomic radius*** *of the atom.* For example, the distance between the nuclei of two carbon atoms in its elemental form is 154 pm, it means its half 77 pm is radius of carbon atom as shown in Figure 3.3:

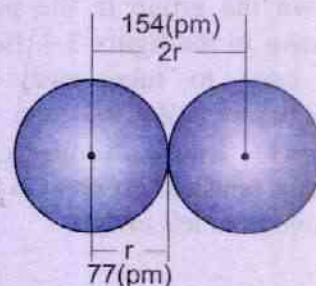



Fig. 3.3 The radius of carbon atom.

When we move from left to right in a period although atomic number increases, yet the size of atoms decreases gradually. It is because with the

increase of atomic number, the effective nuclear charge increases gradually because of addition of more and more protons in the nucleus. But on the other hand addition of electrons takes place in the same valence shell i.e. shells do not increase. There is gradual increase of effective nuclear charge which increases due to addition of protons. This force pulls down or contracts the outermost shell towards the nucleus. For example, atomic size in period 2 decreases from Li (152 pm) to Ne (69 pm).

2nd period elements	^{3}Li	^{4}Be	^{5}B	^{6}C	^{7}N	^{8}O	^{9}F	^{10}Ne
Atomic radii (pm)	152	113	88	77	75	73	71	69

1st group elements	Atomic radii (pm)
^{3}Li	152
^{11}Na	186
^{19}K	227
^{37}Rb	248
^{55}Cs	265

The size of atoms or their radii increases from top to bottom in a group. It is because a new shell of electrons is added up in the successive period, which decreases the effective nuclear charge.

The trend of atomic size of transition elements has slight variation when we consider this series in a period. The atomic size of the elements first reduces or atom contracts and then there is increase in it when we move from left to right in 4th period.

3.2.2 Shielding Effect

The electrons present between the nucleus and the outer most shell of an atom, reduce the nuclear charge felt by the electrons present in the outer most shell. The attractions of outer electrons towards nucleus is partially reduced because of presence of inner electrons. As a result *valance electron experiences less nuclear charge than that of the actual charge, which is called* ***effective nuclear charge*** (Z_{eff}). It means that *the electrons present in the inner shells screen or shield the force of attraction of nucleus felt by the valence shell electrons.* This is called **shielding effect**. With increase of atomic number, the number of electrons in an atom also increases, that results in increase of shielding effect.

The shielding effect increases down the group in the periodic table as shown in the figure 3.4. Because of this it is easy to take away electron from Potassium (**Z**=19) than from Sodium (**Z**=11) atoms. Similarly the shielding effect remains the same in a period if we move from left to right.

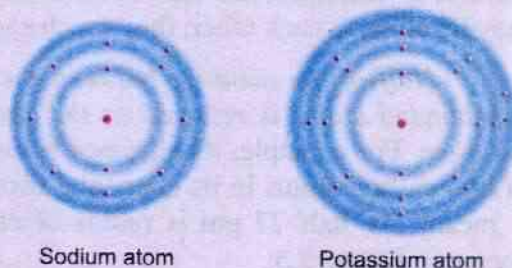



Fig. 3.4: Shielding effect is more in potassium atom than that of sodium atom.

3.2.3 Ionization Energy

The **ionization energy** is *the amount of energy required to remove the most loosely bound electron from the valence shell of an isolated gaseous atom.* The amount of energy needed to remove successive electrons present in an atom increases. If there is only 1 electron in the valence shell, the energy required to remove it will be called first ionization energy. For example, the first ionization energy of sodium atom is $+496 \text{ kJmol}^{-1}$.

$$Na \longrightarrow Na^{+} + e^{-} \qquad \Delta H = +496 \text{ kJmol}^{-1}.$$

But when there are more than one electrons in the valence shell, they can be removed one by one by providing more and more energy. Such as group 2 and 3 elements have more than one electrons in their shells. Therefore, they will have more than one ionization energy values.

If we move from left to right in a period, the value of ionization energy increases. It is because the size of atoms reduces and valence electrons are held strongly by the electrostatic force of nucleus. Therefore, elements on left side of the periodic table have low ionization energies as compared to those on right side of the periodic table as shown for the 2nd period.

2nd period elements	^{3}Li	^{4}Be	^{5}B	^{6}C	^{7}N	^{8}O	^{9}F	^{10}Ne
Ionization energy (kJmol^{-1})	520	899	801	1086	1402	1314	1681	2081

Ionization energy increasing in a period

As we move down the group more and more shells lie between the valence shell and the nucleus of the atom, these additional shells reduce the electrostatic force felt by the electrons present in the outermost shell. Resultantly the valence shell electrons can be taken away easily. Therefore, ionization energy of elements decreases from top to bottom in a group.

1st group elements	Ionization energy (kJmol^{-1})
^{3}Li	520
^{11}Na	496
^{19}K	419
^{37}Rb	403
^{55}Cs	377

Ionization energy decreasing in a group

3.2.4 Electron Affinity

Electron Affinity *is defined as the amount of energy released when an electron is added in the outermost shell of an isolated gaseous atom.*

$$F + e^- \rightarrow F^- \quad \Delta H = -328\ kJmol^{-1}$$

Affinity means attraction. Therefore, electron affinity means tendency of an atom to accept an electron to form an anion. For example, the electron affinity of fluorine is -328 kJ mol^{-1} i.e. one mole atom of fluorine release 328 kJ of energy to form one mole of fluoride ions.

Let us discuss the trend of electron affinity in the periodic table. Electron affinity values increase from left to right in the period.

2nd period elements	^{3}Li	^{4}Be	^{5}B	^{6}C	^{7}N	^{8}O	^{9}F	^{10}Ne
Electron affinity (kJmol^{-1})	−60	>0	−29	−122	0	−141	−328	0

Electron affinity increasing in a period

The reason for this increase is, as the size of atoms decreases in a period, the attraction of the nucleus for the incoming electron increases. That means more is attraction for the electron, more energy will be released.

In a group electron affinity values decrease from top to bottom because the size of atoms increases down the group. With the increase in size of atom shielding effect increases that results in poor attraction for the incoming electron i.e. less energy is released out. For example, as the size of iodine atom is bigger than chlorine, its electron affinity is less than chlorine, as given in the adjacent table.

17th group elements	Electron affinity (kJmol^{-1})
^{9}F	−328
^{17}Cl	−349
^{35}Br	−325
^{53}I	−295

Electron affinity decreasing in a group

3.2.5 Electronegativity

The ability of an atom to attract the shared pair of electrons towards itself in a molecule, is called **electronegativity**. It is an important property especially when covalent type of bonding of elements is under consideration.

The trend of electronegativity is same as of ionization energy and electron affinity. It increases in a period from left to right because higher Z_{eff} shortens distance from the nucleus of the shared pair of electrons. This enhances the power to attract the shared pair of electrons. For example, electronegativity values of period 2 are as follow:

2nd period elements	^{3}Li	^{4}Be	^{5}B	^{6}C	^{7}N	^{8}O	^{9}F
Electronegativity	1.0	1.6	2.0	2.6	3.0	3.4	4.0

Electronegativity increasing in a period

It generally decreases down a group because size of the atom increases. Thus attraction for the shared pair of electrons weakens. For example, electronegativity values of group 17 (halogens) are presented here.

17th group elements	Electro negativity
^{9}F	4.0
^{17}Cl	3.2
^{35}Br	3.0
^{53}I	2.7

Test yourself 3.3

i. Define atomic radius?
ii. What are SI units of atomic radius?
iii. Why the size of atoms decreases in a period?
iv. Define ionization energy.
v. Why the 2nd ionization energy of an elements is higher than first one?
vi. What is the trend of ionization energy in a group?
vii. Why the ionization energy of sodium is less than that of magnesium?
viii. Why is it difficult to remove an electron from halogens?
ix. What is shielding effect?
x. How does shielding effect decrease the forces of electrostatic attractions between nucleus and outer most electrons?
xi. Why does the bigger size atoms have more shielding effect?
xii. Why does the trend of electron affinity and electronegativity is same in a period?
xiii. Which element has the highest electronegativity?

Key Points

- In nineteenth century attempts were made to arrange elements in a systematic manner.
- Dobereiner arranged elements in a group of three called triads.
- Newlands arranged elements in groups of eight like musical notes.

- Mendeleev constructed Periodic Table containing periods and columns, by arranging elements in order of increasing atomic weights.
- There are total eighteen groups and seven periods in the modern Periodic Table.
- Depending on outermost electrons and electronic configuration, element in periodic table are grouped in *s, p, d* and *f* blocks.
- Atomic size increases down a group but decreases along the period .
- Ionization energy decreases down a group but increases along a period.
- Shielding effect is greater in atoms with greater number of electrons.
- Electronegativity increases along a period and decreases down the group.

EXERCISE

Multiple Choice Questions

Put a (✓) on the correct answer

1. **The atomic radii of the elements in Periodic Table:**
   (a) increase from left to right in a period
   (b) increase from top to bottom in a group
   (c) do not change from left to right in a period
   (d) decrease from top to bottom in a group
2. **The amount of energy given out when an electron is added to an atom is called:**
   (a) lattice energy (b) ionization energy
   (b) electronegativity (d) electron affinity
3. **Mendeleev Periodic Table was based upon the:**
   (a) electronic configuration (b) atomic mass
   (c) atomic number (d) completion of a subshell
4. **Long form of Periodic Table is constructed on the basis of:**
   (a) Mendeleev Postulate (b) atomic number
   (c) atomic mass (d) mass number
5. **4^{th} and 5^{th} period of the long form of Periodic Table are called:**
   (a) short periods (b) normal periods
   (c) long periods (d) very long periods
6. **Which one of the following halogen has lowest electronegativity?**
   (a) fluorine (b) chlorine
   (c) bromine (d) iodine
7. **Along the period, which one of the following decreases:**
   (a) atomic radius (b) ionization energy
   (c) electron affinity (d) electronegativity

8. **Transition elements are:**
   (a) all gases (b) all metals
   (c) all non-metals (d) all metalloids

9. **Mark the incorrect statement about ionization energy:**
   (a) it is measured in $kJmol^{-1}$ (b) it is absorption of energy
   (c) it decreases in a period (d) it decreases in a group

10. **Point out the incorrect statement about electron affinity:**
    (a) it is measured in $kJmol^{-1}$ (b) it involves release of energy
    (c) it decreases in a period (d) it decreases in a group

Short Answer Questions

1. Why are noble gases not reactive?
2. Why Cesium (at.no.55) requires little energy to release its one electron present in the outermost shell?
3. How is periodicity of properties dependent upon number of protons in an atom?
4. Why shielding effect of electrons makes cation formation easy?
5. What is the difference between Mendeleev's periodic law and modern periodic law?
6. What do you mean by groups and periods in the Periodic Table?
7. Why and how are elements arranged in 4^{th} period?
8. Why the size of atom does not decrease regularly in a period?
9. Give the trend of ionization energy in a period.

Long Answer Questions

1. Explain the contributions of Mendeleev for the arrangement of elements in his Periodic Table.
2. Show why in a '*period*' the size of an atom decreases if one moves from left to right.
3. Describe the trends of electronegativity in a period and in a group.
4. Discuss the important features of modern Periodic Table.
5. What do you mean by blocks in a periodic table and why elements were placed in blocks?
6. Discuss in detail the periods in Periodic Table?
7. Why and how elements are arranged in a Periodic Table?
8. What is ionization energy? Describe its trend in the Periodic Table?
9. Define electron affinity, why it increases in a period and decreases in a group in the Periodic Table.
10. Justify the statement, bigger size atoms have more shielding effect thus low ionization energy.

Structure of Molecules

Major Concepts

Time allocation

Teaching periods	16
Assessment periods	04
Weightage	8%

Students Learning Outcomes

Students will be able to:

- Find the number of valence electrons in an atom using the Periodic Table.
- Describe the importance of noble gas electronic configurations.
- State the octet and duplet rule.
- Explain how elements attain stability.
- Describe the ways in which bonds may be formed.
- State the importance of electronic configurations in formation of ion.
- Describe formation of cations from an atom of a metallic element.
- Describe formation of anion from a non-metallic element.
- Describe characteristic of ionic bond.
- Recognize a compound as having ionic bonds.
- Identify characteristics of ionic compounds.
- Describe formation of covalent bond between two non-metallic elements.
- Describe with examples single, double and triple covalent bonds.
- Draw electron cross and dot structure of simple covalent molecules containing single, double and triple covalent bonds.

Introduction

The things around us are composed of matter. All matter is made up of the building units 'atoms'. These atoms combine to form molecules, which appear in different states of matter around us. The forces responsible for binding the atoms together in a molecule are called chemical forces or chemical bonds. These bonding forces which keep the atom together will be discussed in this chapter.

4.1 WHY DO ATOMS FORM CHEMICAL BONDS?

It is a universal rule that everything in this world tends to become more stable. Atoms achieve stability by attaining electronic configuration of noble gases (He, Ne or Ar etc) i.e $ns^2 np^6$. Having 2 or 8 electrons in the valence shell is sign of stability. *Attaining two electrons in the valence shell is called **duplet** rule while attaining eight electrons in the valence shell is called **octet** rule.*

The noble gases do have 2 or 8 electrons in their valence shells. It means all the noble gases have their valence shells completely filled. Their atoms do not have vacant space in their valence shell to accommodate extra electrons. Therefore, noble gases do not gain, lose or share electrons. That is why they are non-reactive.

The importance of the noble gas electronic configuration lies in the fact that all other atoms try their best to have the noble gas electronic configuration. For this purpose atoms combine with one another, which is called chemical bonding. In other words, atoms form chemical bonds to achieve stability by acquiring inert gas electron configuration.

An atom can accommodate 8 electrons in its valence shell in three ways:

i. By giving valence shell electrons (if they are less than three) to other atoms.

ii. By gaining electrons from other atoms (if the valence shell has five or more electrons in it).

iii. By sharing valence electrons with other atoms.

It means every atom has a natural tendency to have 2 or 8 electrons in its valence shell. The atoms having less than 2 or 8 electrons in their valence shells are unstable.

Now the question arises that how can we identify the way an atom reacts? The position of an atom in the periodic table indicates its group number. As we have studied in chapter 3, the group number is assigned on the basis of valence shell electrons. For example, group 1 has only 1 electron in its valence shell and group 17 has 7 electrons in its valence shell. Mode of reaction of an atom depends upon its number of valence shell electrons. It is discussed in the next sections.

4.2 CHEMICAL BOND

A **chemical bond** *is defined as a force of attraction between atoms that holds them together in a substance.* In other words during bond formation there is some force which holds the atoms together.

This attaining of 8 electrons configuration in the outermost shell either by sharing, by losing or by gaining electrons, is called **octet rule**. This octet rule only symbolizes that noble gas electronic configuration should be attained by atoms when they combine or react. For elements like hydrogen or helium; which have only

s-subshell, this becomes '*duplet rule*'. It plays a significant role in understanding the formation of chemical bond between atoms.

If the bond formation is between ions, it is due to an electrostatic force of attraction between them. But if bond formation is between similar atoms or between the atoms that have comparable electronegativities, then the chemical bond formation is by '*sharing*' of electrons. This sharing of electrons may be mutual or one sided.

When two approaching atoms come closer, the attractive as well as repulsive forces become operative. The formation of a chemical bond is a result of net attractive forces which dominate. The energy of that system is lowered and molecule is formed. Otherwise if repulsive forces become dominant no chemical bond will be formed. In that case there will be increase in the energy of the system due to creation of repulsive forces.

4.3 TYPES OF CHEMICAL BOND

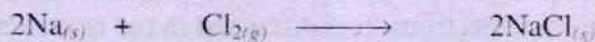
The valence electrons, which are involved in chemical bonding, are termed as **bonding electrons**. They usually reside in the incomplete or partially filled outer most shell of an atom. Depending upon the way how these valence electrons are involved in bonding, they result in following four types of chemical bonds:

- Ionic Bond
- Covalent Bond
- Dative Covalent or Coordinate Covalent Bond
- Metallic Bond

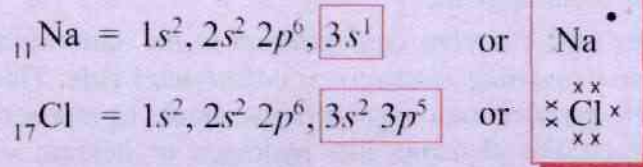
4.3.1 Ionic Bond

The elements of Group-1 and Group-2 being metals have the tendency to lose their valence electrons forming positively charged ions. Whereas non-metals of Group-15 to Group-17 have the tendency to gain or accept electrons. They are electronegative elements with high electron affinities. If atoms belonging to these two different groups, metals and non-metals, are allowed to react, chemical bond is formed. *This type of chemical bond, which is formed due to complete transfer of electron from one atom to another atom, is called* **ionic bond**.

The formation of NaCl is a good example of this type of bond.

$$2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$$

Sodium chloride is a simple compound formed by sodium (**Z** =11) and Chlorine (**Z**=17) atoms. The ground state electronic configuration of these elements is shown below:

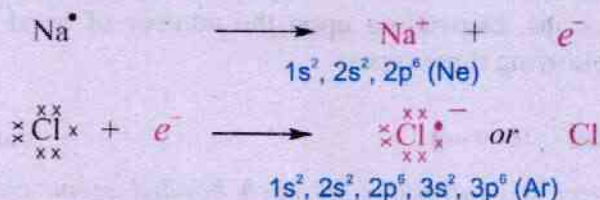
$${}_{11}Na = 1s^2, 2s^2\ 2p^6, \boxed{3s^1} \quad \text{or} \quad Na^{\bullet}$$

$${}_{17}Cl = 1s^2, 2s^2\ 2p^6, \boxed{3s^2\ 3p^5} \quad \text{or} \quad \overset{\times\times}{\underset{\times\times}{\times Cl \times}}$$

The frames indicate electrons in the valence shells of these elements: sodium has only one electron and chlorine has seven electrons. Sodium being electropositive element has the tendency to lose electron and chlorine being an electronegative element has the tendency to gain electron. Therefore, they form positive and negative ions by losing and gaining electrons respectively. They attain electronic configuration to the nearest noble gases.

$$Na^{\bullet} \longrightarrow Na^{+} + e^{-}$$

$1s^2, 2s^2, 2p^6$ (Ne)

$$\overset{\times\times}{\underset{\times\times}{\times Cl \times}} + e^{-} \longrightarrow \left[\overset{\times\times}{\underset{\times\times}{\times Cl \bullet}}\right]^{-} \quad or \quad Cl^{-}$$

$1s^2, 2s^2, 2p^6, 3s^2, 3p^6$ (Ar)

By losing one electron from the outermost shell, sodium becomes Na^+ ion and it is left with 8 electrons in the second shell which will now become the valence shell. By gaining one electron, chlorine atom now also has eight electrons in its outermost shell and becomes Cl^- ion. Both of these atoms are now changed into oppositely charged ions. They stabilize themselves by combining with each other due to electrostatic force of attraction between them such as:

$$Na^{+} + Cl^{-} \longrightarrow NaCl$$

It is to be noted that only valence shell electrons take part in this type of bonding, while other electrons are not involved. In such type of reaction heat is usually given out. The compounds formed due to this type of bonding are called **ionic compounds**.

Test yourself 4.1

i. *Why does sodium form a chemical bond with chlorine?*
ii. *Why does sodium lose an electron and attains +1 charge?*
iii. *How do atoms follow octet rule?*
iv. *Which electrons are involved in chemical bonding?*
v. *Why does group 1 elements prefer to combine with group 17 elements.*
vi. *Why chlorine can accept only 1 electron?*

4.3.2 Covalent Bond

The elements of Group-13 to Group-17 when allowed to react with each other, they form a chemical bond by mutual sharing of their valence shell electrons. This type of *bond, which is formed due to mutual sharing of electrons,* is called a **covalent bond.**

The energy changes during the covalent bond formation are of considerable value. When two atoms approach each other attractive forces develop between electrons of one atom and nucleus of the other atom. Simultaneously repulsive forces between electrons of the two atoms as well as between their nuclei are also created. When the attractive forces dominate due to decrease in distance between those two

atoms, a chemical bond is formed between them. The formation of hydrogen, chlorine, nitrogen and oxygen gases are few examples of this type of bonding.

Types of covalent bonds

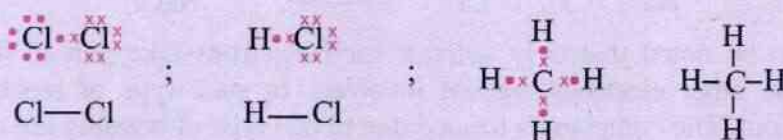
As described above, the covalent bond is formed by mutual sharing of electrons between two atoms. *The electrons that pair up to form a chemical bond are called* '**bond pair**' electrons. Depending upon the number of bond pairs, covalent bond is classified into following three types:

Single Covalent bond (—)

When one electron is contributed by each bonded atom, one bond pair is formed and it forms a **single covalent bond**. While drawing the structure of such molecules the single bond pair is indicated by a line between those two atoms. A few examples of molecules with single covalent bonds are hydrogen (H_2), chlorine (Cl_2), hydrochloric acid (HCl) and methane (CH_4).

$$H\cdot + {}_{\times}H \longrightarrow H \cdot\times H \quad \text{or} \quad H—H \ ; \quad H_2$$

single covalent bond

Cl—Cl ; H—Cl ; H−C−H (with H above and below C)

Double Covalent bond (═)

When each bonded atom contributes two electrons, two bond pairs are shared and a **double covalent bond** is formed. These bond pairs are indicated as double line between those atoms in the structure of such molecules. The molecules like oxygen (O_2) gas and ethene (C_2H_4) show such type of double covalent bonds.

$$O + O \longrightarrow O \,::\, O \quad \text{or} \quad O{=}O \ ; \quad O_2$$

double covalent bond

H₂C=CH₂ (ethene: C=C with two H on each C)

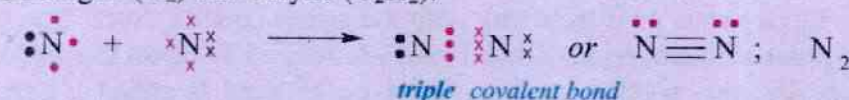
Triple Covalent Bond (≡)

When each bonded atom contributes three electrons, three bond pairs are involved in bond formation. This type is called **triple covalent bond**. Three small lines are used to indicate these three pairs of electrons between those atoms in the

molecules of such compounds. The examples of molecules having triple covalent bonds are nitrogen (N_2) and ethyne (C_2H_2).

$$:\dot{N}\cdot + \times N \times \longrightarrow :N⋮⋮N\times \quad or \quad \ddot{N}\equiv\ddot{N}\ ; \quad N_2$$

triple covalent bond

$$H\cdot\times C⋮⋮C\cdot\times H \qquad H-C\equiv C-H$$

By this mutual sharing of valence shell electrons, each of the contributing atom attains the '*Octet*' or nearest noble gas electronic configuration.

The electronic configuration of the valence shells of atoms is shown in small 'dots' or 'crosses' around the symbol of the element. Each dot or cross represents an electron. This is a standard method of Lewis to describe the electronic configuration of valence shell of an atom. It is called ***Lewis Structure Diagram.***

4.3.3 Dative Covalent or Coordinate Covalent Bond

Coordinate covalent or **dative covalent bonding** is a type of *covalent bonding in which the bond pair of electrons is donated by one bonded atom only*. The atom which donates the electron pair is called **donor** and the atom which accepts the electron pair is called **acceptor.** A small arrow is usually used to indicate the atom and pair of electron being donated. The head of arrow is towards the acceptor atom.

The non-bonded electron pair available on an atom, like the one available on nitrogen in ammonia,($\ddot{N}H_3$*) is called* a **lone pair**. When a proton (H^+) approaches a molecule with a *lone pair* of electrons, that lone pair is donated to H^+ and a **coordinate covalent bond** is formed, e.g. formation of ammonium radical (NH_4^+).

In the formation of BF_3 (boron trifluoride) molecule, three valence electrons of boron atom (Z= 5) pair up with three electrons, one from each three fluorine atoms. The boron atom even after this sharing of electrons (covalent bond formation), remains short or deficit of two electrons in its outermost shell. Now if a molecule with a lone pair approaches this molecule, it accepts lone pair from that *donor* and forms a coordinate covalent bond. The lone pair on nitrogen of ammonia molecule makes it a good donor molecule to form a coordinate covalent bond as shown in figure 4.1.

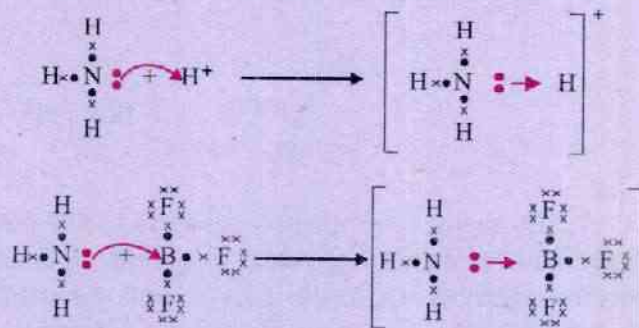



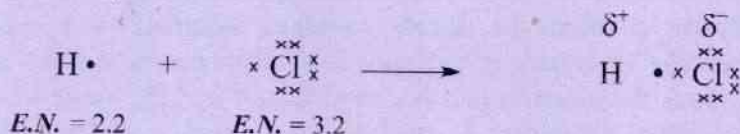
Fig. 4.1 Dative covalent bond (red arrow)

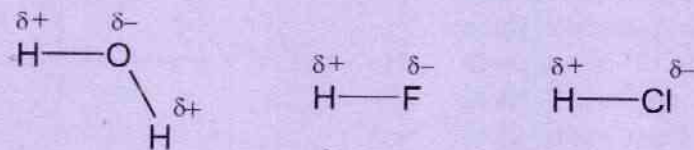
4.3.4 Polar and Non-polar Covalent Bond

If a covalent bond is formed between two similar atoms (homo-atoms), the shared pair of electrons is attracted by both the atoms equally. Such type of bond is called **non-polar** covalent bond. These bonds are formed by equal sharing of electron pair between the two bonding atoms. This type of bond is called a pure covalent bond. For example bond formation in H_2 and Cl_2.

If the covalent bond is formed between two different types of atoms (hetro-atoms) then the bond pair of electrons will not be attracted equally by the bonded atoms. One of the atoms will attract the bond pair of electrons more strongly than the other one. This atom(element) will be called as more electronegative.

When there is difference of electronegativity between two covalently bonded atoms, there will be unequal attraction for the bond pair of electrons between such atoms. It will result in the formation of polar covalent bond. The difference between electronegativities of hydrogen and chlorine is 1.0. As the electronegativity of chlorine is more, it attracts the shared pair of electron towards itself with a greater force. A partial negative charge is therefore created on chlorine and in turn a partial positive charge on hydrogen due to electronegativity difference. It creates polarity in the bond and is called polar covalent bond.

$$\underset{E.N. = 2.2}{H\cdot} \quad + \quad \underset{E.N. = 3.2}{\times \overset{\times\times}{\underset{\times\times}{Cl}} \overset{\times}{\times}} \longrightarrow \overset{\delta^+}{H} \;\bullet\times \overset{\delta^-}{\overset{\times\times}{\underset{\times\times}{Cl}}} \overset{\times}{\times}$$

The delta (δ) sign indicates partial positive or partial negative charge that is developed due to unequal sharing of shared pair or bonded pair of electrons. The compounds resulting from polar covalent bonds are called **polar compounds**. For example water, hydrogen fluoride and hydrogen chloride.

$$\overset{\delta+}{H}-\overset{\delta-}{O}-\overset{\delta+}{H} \qquad \overset{\delta+}{H}-\overset{\delta-}{F} \qquad \overset{\delta+}{H}-\overset{\delta-}{Cl}$$

By using electronegativity values, it is possible to predict whether a chemical bond will be ionic or covalent in nature. A bond formed between elements of high electronegativity (halogen group) and elements of low electronegativity (alkali metals) are ionic in nature. There is complete transfer of electrons between them. The bond between elements of comparable electronegativities will be covalent in nature as the bond between carbon and hydrogen in methane, or nitrogen and hydrogen in ammonia. *If the difference of electronegativities between two elements is more than 1.7 the bond between them will be predominantly ionic bond and if it is less than 1.7, the bond between two atoms will be predominantly covalent.*

Test yourself 4.2

i. *Give the electronic configuration of carbon atom.*
ii. *What type of elements have tendency of sharing of electrons?*
iii. *If repulsive forces dominate to attractive forces will a covalent bond form?*
iv. *Considering the electronic configuration of nitrogen atom, how many electrons are involved in bond formation and what type of covalent bond is formed.*
v. *Point out the type of covalent bonds in the following molecules*
CH_4, C_2H_4, H_2, N_2, and O_2
vi. *What is a lone pair? How many lone pairs of electrons are present on nitrogen in ammonia?*
vii. *Why is the BF_3 electron deficient?*
viii. *What types of electron pairs make a molecule good donor?*
ix. *What is difference between bonded and lone pair of electron and how many bonded pair of electrons are present in NH_3 molecule?*
x. *What do you mean by delta sign and why it develops?*
xi. *Why does oxygen molecule not form a polar covalent bond?*
xii. *Why has water polar covalent bonds?*

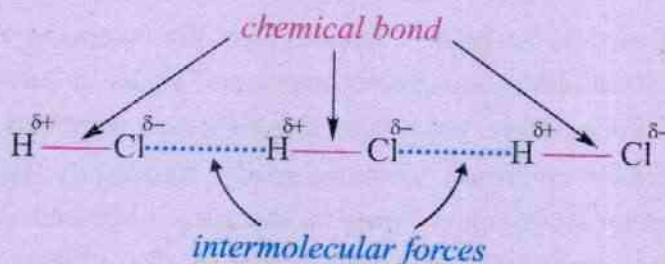
4.3.5 Metallic Bond

The **metallic bond** *is defined as a bond formed between metal atoms (positively charged ions) due to mobile or free electrons.* The different properties shown by metals such as high melting and boiling points, good conductions of heat and electricity, hard and heavy nature, suggest existence of different type of chemical bond between atoms of metals.

In case of metals, the hold of nucleus over the outermost electrons is weak because of large sized atoms and greater number of shells in between nucleus and valence electrons. Furthermore, because of low ionization potentials, metals have the tendency to lose their outermost electrons easily. Resultantly, these loose or free electrons of all metal atoms move freely in the spaces between atoms of a metal. None of these electrons is attached to any particular atom. Either they belong to a common pool, or belong to all the atoms of that metal. Nuclei of metal atoms appear submerged in sea of these free mobile electrons. These mobile electrons are responsible for holding the atoms of metals together forming a metallic bond. A simple metallic bond is shown in figure 4.2.

Fig. 4.2 A schematic diagram of Copper wire showing its positive nuclei (+) embedded in sea of free electrons (o) making 'Metallic Bonding'

4.4 INTERMOLECULAR FORCES

As discussed earlier, the forces that hold atoms in a compound are chemical bonds. In addition to these strong bonding forces, relatively weak forces also exist in between the molecules, which are called **intermolecular forces**. The bonding and intermolecular forces of hydrochloric acid are shown below:

It requires about 17 *kJ* energy to break these *intermolecular forces* between one *mole* of liquid hydrogen chloride molecules to convert it into gas. Whereas, about 430 *kJ* energy is required to break the *chemical bond* between hydrogen and chlorine atoms in 1 *mole* of hydrogen chloride.

4.4.1 Dipole - Dipole Interaction

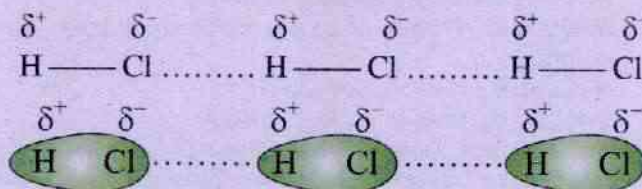
All intermolecular forces, which are collectively called **van der Waals** forces, are electrical in nature. They result from the attractions of opposite charges which may be temporary or permanent. The unequal sharing of electrons between two different types of atoms make one end of molecule slightly positive and other end slightly negatively charged. As shared pair of electron is drawn towards more

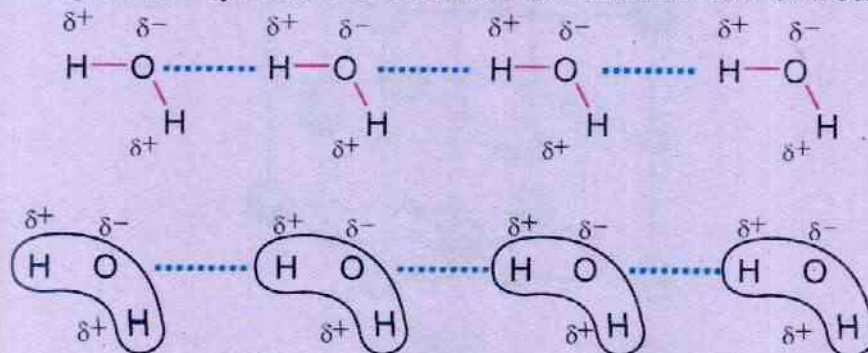
electronegative atom, it is partially negatively charged, as chlorine in hydrogen chloride. The other end automatically becomes partially positively charged.

$$\overset{\delta^+}{H} — \overset{\delta^-}{Cl}$$

When partial positive and partial negative charges exist at different positions in a molecule, the adjacent molecules will arrange themselves in such a way that negative end of that molecule comes near to positive end of other molecule. It results in a net forces of attraction between oppositely charged ends of two adjacent molecules. These attractive forces are called dipole – dipole interactions as represented in HCl:

$$\overset{\delta^+}{H} — \overset{\delta^-}{Cl} \cdots\cdots \overset{\delta^+}{H} — \overset{\delta^-}{Cl} \cdots\cdots \overset{\delta^+}{H} — \overset{\delta^-}{Cl}$$

$$\overset{\delta^+}{H}\ \overset{\delta^-}{Cl} \cdots\cdots \overset{\delta^+}{H}\ \overset{\delta^-}{Cl} \cdots\cdots \overset{\delta^+}{H}\ \overset{\delta^-}{Cl}$$

4.4.2 Hydrogen Bonding

Hydrogen bonding is a special type of intermolecular forces present in the permanently polar molecules. This bonding can be considered unique dipole-dipole attraction. This force of attraction develops between molecules that have a hydrogen atom bonded to a small, highly electronegative atom with lone pairs of electrons such as nitrogen, oxygen and fluorine. The covalent bond between hydrogen atom and other atom becomes polar enough to create a partial positive charge on hydrogen atom and a partial negative charge on the other atom. The small size and high partial positive charge on the hydrogen atom enables it to attract highly electronegative (N,O or F) atom of the other molecule. *So, partially positively charged hydrogen atom of one molecule attracts and forms a bond with the partially negatively charged atom of the other molecule, the bonding is called* **hydrogen bonding**. This force of attraction is represented by a dotted line between the molecules as shown below:

Hydrogen bonding affects the physical properties of the molecules. Due to this boiling points of the compounds are affected greatly. For example, boiling point of

water (100 °C) is higher than that of alcohol (78 °C) because of more and stronger hydrogen bonding in water.

The important phenomenon of floating of ice over water is because of hydrogen bonding. The density of ice at 0 °C (0.917 gcm^{-3}) is less than that of liquid water at 0 °C (1.00 gcm^{-3}). In the liquid state water molecules move randomly. However, when water freezes, the molecules arrange themselves in an ordered form, that gives them open structure. This process expands the molecules, that results in ice being less dense as compared to water.

Test yourself 4.3

i. What type of elements form metallic bonds?
ii. Why is the hold of nucleus over the outermost electrons in metals weak?
iii. Why the electrons move freely in metals?
iv. Which types of electrons are responsible for holding the atoms together in metals.
v. Why a dipole develops in a molecule?
vi. What do you mean by induced dipole?
vii.. Why are dipole forces of attraction not found in halogen molecules?
viii What types of attractive forces exist between HCl molecules?
ix. Define intermolecular forces; show these forces among HCl molecule.

4.5. NATURE OF BONDING AND PROPERTIES

Properties of the compounds depend upon the nature of bonding present in them. Let us discuss the effects of nature of bonding on the properties of compounds.

4.5.1 Ionic Compounds

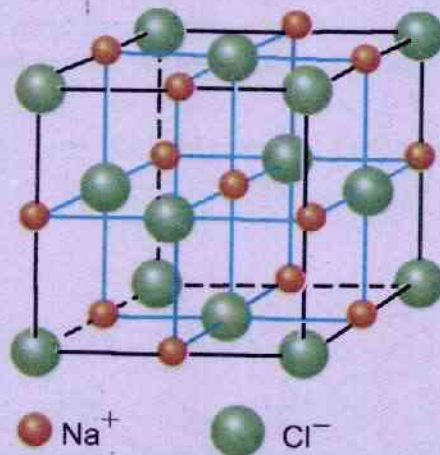
Ionic compounds are made up of positively and negatively charged ions. Thus they consist of ions and not the molecules. These positively and negatively charged ions are held together in a solid or crystal form with strong electrostatic attractive forces. The orderly arrangement of Na^+ and Cl^- ions in a solid crystal of sodium chloride is shown in figure 4.3.

Na^+ Cl^-

Figure 4.3 Regular arrangement of Na^+ and Cl^- ions in solid crystal of NaCl

The ionic compounds have following properties:

i. Ionic compounds are mostly crystalline solids.

ii. Ionic compounds in solid state have negligible electrical conductance but they are good conductors in solution and in the molten form. It is due to presence of free ions in them.

iii. Ionic compounds have high melting and boiling points. For example, sodium chloride has melting point 800 °C and a boiling point 1413 °C. As ionic compounds are made up of positive and negative ions, there exist strong electrostatic forces of attraction between oppositely charged ions. So, a great amount of energy is required to break these forces.

iv. They dissolve easily in polar solvents like water. Water has high dielectric constant that weakens the attraction between ions.

4.5.2 Covalent Compounds

The covalent compounds are made up of molecules that are formed by mutual sharing of electrons between their atoms i.e. covalent bonds. A covalent bond is generally regarded as weaker than an ionic bond. Covalent compounds are made up of two or more non-metals, e.g. H_2, CH_4, CO_2, H_2SO_4, $C_6H_{12}O_6$. Lower molecular mass covalent compounds are gases or low boiling liquids. Contrary to it, higher molecular mass covalent compounds are solids. General properties shown by covalent compound are as follows:

i. They have usually low melting and boiling points.

ii. They are usually bad conductors of electricity. The compounds having polar character in their bonding are conductor of electricity when they dissolve in polar solvents.

iii. They are usually insoluble in water but are soluble in non-aqueous solvents like benzene, ether, alcohol and acetone.

iv. Large molecules with three dimensional bonding form covalent crystals which are very stable and hard. They have very high melting and boiling points.

Polar and Non-Polar Compounds

As discussed earlier the polarity in a chemical bond is due to difference in electronegativities of the bonding atoms. On the **Pauling Scale** fluorine has been given an electronegativity value of 4.0. The values for other elements are calculated relative to it.

Properties of non-polar and polar covalent compounds differ to some extent. Non-polar covalent compounds usually do not dissolve in water while polar covalent compounds usually dissolve in water. Similarly non-polar compounds do not conduct electricity but an aqueous solution of a polar compound usually conduct electricity due to the formation of ions as a result of its reaction with water.

4.5.3 Coordinate Covalent Compounds

Their properties are mostly similar to those of covalent compounds. As the nuclei in these compounds are held by shared pair of electrons, therefore, they do not form ions in water. Due to their covalent nature they form solutions in organic solvents and are very less soluble in water. Usually they are rigid compounds with a dipole.

4.5.4 Metals

Metals have common property of conducting heat and electricity. It gives them prime role in many industires. Major properties shown by the metals are as follows:

i. They show metallic luster.

ii. They are usually malleable and ductile. Malleability is the property by virtue of which a metal can be rolled into sheets, while ductility is the property by virtue of which a metal can be drawn into wires.

iii. They have usually high melting and boiling points.

iv. Being greater in size they have low ionization energies and form cations (M^+) very easily.

v. They are good conductors of heat and electricity in solid and liquid state due to mobile electrons.

Test yourself 4.4

i. Why the ionic compounds have high melting and boiling points?
ii. What do you mean by malleability?
iii. Why are ionic compounds easily soluble in water?
iv. What type of bond exists in sodium chloride ?
v. Why the covalent compounds of bigger size molecules have high melting points?
vi. (a): What is the electronegativity difference between the following pair of elements (atoms). Predict the nature of the bond between them?
(a) H and Cl (b) H and Na (c) Na and I (d) K and Cl
(b): Comparing the electronegativity differences, arrange these compounds in increasing ionic strength.

Synthetic Adhesives

Although natural adhesives are less expensive to produce, but most important adhesives used now a days are synthetic. Adhesives based on synthetic resins and rubbers excel in versatility and performance. Synthetic adhesives can be produced in a sufficient supply with uniform properties and they can be modified in many ways. The polymers or resins used in synthetic adhesives fall into two general categories—thermoplastics and thermoseting. One form of polymer used industrially is epoxy adhesive.

AIR CRAFTS, CARS, TRUCKS AND BOATS ARE PARTIALLY HELD TOGETHER WITH EPOXY ADHESIVES

Epoxy is polymer that is formed from two different chemicals. These are referred to as resin and the hardener. Epoxy adhesives are called structural adhesives. These high-performance adhesives are used in the construction of aircraft, automobiles, bicycles, boats, golf clubs, where high strength bonds are required. Epoxy adhesives can be developed to suit almost any application. They can be made flexible or rigid, transparent or opaque even colored as well as fast or slow setting. Epoxy adhesives are good heat and chemical resistant. Because of these properties, they are given the name of engineering adhesives.

KEY POINTS

- Atoms of different elements react to attain noble gas configuration, which is stable one.
- Chemical bonds may be formed by complete transfer of electrons (ionic); mutual sharing (covalent) or by donation from an atom(coordinate or dative covalent).
- Metals have the tendency to lose electrons easily forming cations.
- Non-metals have tendency to gain electrons and form anions.
- In ionic bonding strong electrostatic force hold ions together.
- Ionic compounds are solids with high melting and boiling points.
- Covalent bonds among non-metals are weaker than ionic bonds.
- Ionic bonds are non-directional, but covalent bonds are formed in a particular direction.
- Covalent bonds formed between similar atoms are non-polar while between different atoms are polar.
- In covalent bonding *single*, *double* or *triple* covalent bond is formed by sharing of one, two or three electron pairs by the bonded atoms.
- Coordinate covalent bond is formed between electron pair donors and electron pair acceptors.
- Metallic bond is formed between metal atoms due to free electrons.
- In addition to chemical bonds, intermolecular forces of attraction exist between polar molecules.
- Hydrogen bonding exists between the hydrogen atom of one molecule and highly electronegative atom of other molecule.
- Hydrogen bonds affect the physical properties of the compounds.

- Properties of the compounds depend upon the nature of bonding present in the compound.
- Ionic compounds are crystalline solid with high melting and boiling points.
- Covalent compounds exist in molecular form in three physical states.
- Polar and non- polar covalent compounds have different properties.
- Metals have shining surface. They are good conductor of electricity and are malleable and ductile

EXERCISE

Multiple Choice Questions

Put a (✓) on the correct answer

1. **Atoms react with each other because:**
   (a) they are attracted to each other. (b) they are short of electrons
   (c) they want to attain stability (d) they want to disperse
2. **An atom having six electrons in its valence shell will achieve noble gas electronic configuration by:**
   (a) gaining one electron (b) losing all electrons
   (c) gaining two electrons (d) losing two electrons
3. **Considering the electronic configuration of atoms which atom with the given atomic number will be the most stable one?**
   (a) 6 (b) 8 (c) 10 (d) 12
4. **Octet rule is:**
   (a) description of eight electrons
   (b) picture of electronic configuration
   (c) pattern of electronic configuration
   (d) attaining of eight electrons
5. **Transfer of electrons between atoms results in:**
   (a) metallic bonding (b) ionic bonding
   (c) covalent bonding (d) coordinate covalent bonding
6. **When an electronegative element combines with an electropositive element the type of bonding is:**
   (a) covalent (b) ionic
   (c) polar covalent (d) coordinate covalent
7. **A bond formed between two non-metals is expected to be:**
   (a) covalent (b) ionic
   (c) coordinate covalent (d) metallic
8. **A bond pair in covalent molecules usually has:**
   (a) one electron (b) two electrons
   (c) three electrons (d) four electrons

9. Which of the following compounds is not directional in its bonding?
(a) CH_4 (b) KBr (c) CO_2 (d). H_2O

10. Ice floats on water because:
(a) ice is denser than water (b) ice is crystalline in nature
(c) water is denser than ice (d) water molecules move randomly

11. Covalent bond involves the
(a) donation of electrons (b) acceptance of electrons
(c) sharing of electrons (d) repulsion of electrons

12. How many covalent bonds does C_2H_2 molecule have?
(a) two (b) three (c) four (d) five

13. Triple covalent bond involves how many electrons?
(a) eight (b) six (c) four (d) only three

14. Which pair of the molecules has same type of covalent bonds?
(a) O_2 and HCl (b) O_2 and N_2
(c) O_2 and C_2H_4 (d) O_2 and C_2H_2

15. Identify the compound which is not soluble in water.
(a) C_6H_6 (b) NaCl (c) KBr (d) $MgCl_2$

16. Which one of the following is an electron deficient molecule?
(a) NH_3 (b) BF_3 (c) N_2 (d) O_2

17. Identify which pair has polar covalent bonds.
(a) O_2 and Cl_2 (b) H_2O and N_2
(c) H_2O and C_2H_2 (d) H_2O and HCl

18. Which one of the following is the weakest force among the atoms?
(a) ionic force (b) metallic force
(c) intermolecular force (d) covalent force

Short Answer Questions

1. Why do atoms react?
2. Why is the bond between an electropositive and an electronegative atom ionic in nature?
3. Ionic compounds are solids. Justify.
4. More electronegative elements can form bonds between themselves. Justify.
5. Metals are good conductor of electricity. Why?
6. Ionic compounds conduct electricity in solution or molten form. Why?
7. What type of covalent bond is formed in nitrogen molecule.
8. Differentiate between lone pair and bond pair of electrons.

9. Describe at least two necessary conditions for the formation of a covalent bond.
10. Why HCl has dipole-dipole forces of attraction?
11. What is a triple covalent bond, explain with an example?
12. What is difference between polar and non-polar covalent bonds, explain with one example of each?
13. Why a covalent bond becomes polar?
14. What is relationship between electronegativity and polarity?
15. Why does ice float on water?
16. Give the characteristic properties of ionic compounds.
17. What characteristic properties do the covalent compound have?

Long Answer Questions

1. What is an ionic bond? Discuss the formation of ionic bond between sodium and chlorine atoms?
2. How can you justify that bond strength in polar covalent compounds is comparable to that of ionic compound?
3. What type of covalent bonds are formed between hydrogen, oxygen and nitrogen? Explain their bonding with dot and cross model.
4. How a covalent bond develops ionic character in it? Explain.
5. Explain the types of covalent bonds with at least one example of each type.
6. How a coordinate covalent bond is formed? Explain with examples?
7. What is metallic bond? Explain the metallic bonding with the help of a diagram.
8. Define hydrogen bonding. Explain that how these forces affect the physical properties of compounds.
9. What are intermolecular forces? Compare these forces with chemical bond forces with reference to HCl molecule?
10. What is a chemical bond and why do atoms form a chemical bond?
11. What is octet rule? Why do atoms always struggle to attain the nearest noble gas electronic configuration?

Physical States of Matter

Major Concepts

Gaseous State

5.1 Typical properties
5.2 Laws related to gases

Liquid State

5.3 Typical Properties

Solid State

5.4 Typical Properties
5.5 Types of Solids
5.6 Allotropy

Time allocation	
Teaching periods	16
Assessment periods	04
Weightage	10%

Students Learning Outcomes

Students will be able to:

- Effect on the volume of a gas by a change in the a. pressure b. temperature.
- Compare the physical states of matter with regard to intermolecular forces present between them.
- Account for pressure-volume changes in a gas using Boyle's Law.
- Account for temperature-volume changes in a gas using Charles' Law.
- Explain the properties of gases(diffusion, effusion and pressure).
- Explain the properties of liquids like evaporation, vapour pressure, boiling point.
- Explain the effect of temperature and external pressure on vapour pressure and boiling point.
- Describe the physical properties of solids (melting and boiling points).
- Differentiate between amorphous and crystalline solids.
- Explain the allotropic forms of solids.

Introduction

Matter exists in three physical states i.e. gas, liquid and solid. The simplest form of matter is the gaseous state. Liquids are less common and most of the matter exists as solid. Matter in gaseous state does not have definite shape and volume. Therefore, gases occupy all the available space. Their intermolecular forces are very weak. Pressure is a significant property of gases. The effect of pressure and temperature on volume of a gas has been studied quite extensively.

The liquid state has strong intermolecular forces hence it has definite volume but it does not have definite shape. It attains the shape of the container in which it is kept. Liquids evaporate and their vapours exert pressure. When vapour pressure of a liquid becomes equal to external pressure, it boils. Liquids are less mobile than gases therefore, they diffuse slowly.

The solid state has definite volume and shape. They are rigid and denser than liquids and gases. They exist in amorphous or crystalline forms.

GASEOUS STATE

5.1 TYPICAL PROPERTIES

Gases have similar physical properties. A few typical properties are discussed here.

5.1.1 Diffusion

Gases can diffuse very rapidly. **Diffusion** *is defined as spontaneous mixing up of molecules by random motion and collisions to form a homogeneous mixture*. Rate of diffusion depends upon the molecular mass of the gases. Lighter gases diffuse rapidly than heavier ones. For example, H_2 diffuses four times faster than O_2 gas .

5.1.2 Effusion

It is escaping of gas molecules through a tiny hole into a space with lesser pressure. For example when a tyre gets punctured, air effuses out. Effusion depends upon molecular masses, lighter gases effuse faster than heavier gases.

5.1.3 Pressure

Gas molecules are always in continuous state of motion. Hence when molecules strike with the walls of the container or any other surface, they exert pressure. **Pressure (*P*)** *is defined as the force(F) exerted per unit surface area*(**A**).

$$P = F / A$$

The ***SI*** unit of force is **Newton** and that of area is m^2. Hence pressure has ***SI*** unit of $N\,m^{-2}$. It is also called **Pascal** (***Pa***) .

$$\text{One Pascal (Pa)} = 1\ \text{Nm}^{-2}$$

Barometer is used to measure atmospheric pressure and manometer is used to measure pressure in the laboratory.

Standard Atmospheric Pressure

It is the pressure exerted by the atmosphere at the sea level. *It is defined as the pressure exerted by a mercury column of 760 mm height at sea level.* It is sufficient pressure to support a column of mercury 760 *mm* in height at sea level.

1 atm = 760 mm of Hg = 760 torr (1 mm of Hg = one torr)

= 101325 Nm^{-2} = 101325 Pa

5.1.4 Compressibility

Gases are highly compressible due to empty spaces between their molecules. When gases are compressed, the molecules come closer to one another and occupy less volume as compared to the volume in uncompressed state.

5.1.6 Mobility

Gas molecules are always in state of continuous motion. They can move from one place to another because gas molecules possess very high kinetic energy. They move through empty spaces that are available for the molecules to move freely. This mobility or random motion results in mixing up of gas molecules to produce a homogeneous mixture.

5.1.7 Density of Gases

Gases have low density than liquids and solids. It is due to light mass and more volume occupied by the gas molecules. Gas density is expressed in *grams* per dm^3. Whereas, liquid and solid densities are expressed in *grams* per cm^3 i.e. liquids and solids are 1000 times denser than gases. The density of gases increases by cooling because their volume decreases. For example, at normal atmospheric pressure, the density of oxygen gas is 1.4 $g\ dm^{-3}$ at 20°C and 1.5 $g\ dm^{-3}$ at 0°C.

Test yourself 5.1

i. Why the rate of diffusion of gases is rapid than that of liquids?
ii. Why are the gases compressible?
iii. What do you mean by Pascal. How many Pascals are equal to 1 atm?
iv. Why the density of a gas increases on cooling?
v. Why is the density of gas measured in g dm^{-3} while that of a liquid in g cm^{-3}?
vi. Convert the following
a. 70 cm Hg to atm
b. 3.5 atm to torr
c. 1.5 atm to Pa

5.2 LAWS RELATED TO GASES

5.2.1 Boyle's Law

In 1662 **Robert Boyle** studied the relationship between the volume and pressure of a gas at constant temperature. He observed that *volume of a given mass of a gas is inversely proportional to its pressure provided the temperature remains constant.*

According to this law the volume (V) of a given mass of a gas decreases with the increase of pressure (P) and vice versa. Mathematically it can be written as:

$$\text{Volume} \propto \frac{1}{\text{Pressure}} \quad \text{or} \quad V \propto \frac{1}{P}$$

Robert Boyle (1627-1691) was natural philosopher, chemist, physicist and inventor. He is famous for 'Boyle's law of gases'.

$$\text{or} \quad V = \frac{k}{P} \quad \text{or} \quad VP = k = \text{constant}$$

Where *'k'* is proportionality constant. The value of ***k*** is same for the same amount of a given gas. Therefore, **Boyle's law** can be *stated as the product of pressure and volume of a fixed mass of a gas is constant at a constant temperature.*

If $P_1V_1 = k$ Then $P_2V_2 = k$

where P_1 = initial pressure P_2 = final pressure

V_1 = initial volume V_2 = final volume

As both equations have same constant therefore, their variables are also equal to each other.

$$\therefore \quad P_1V_1 = P_2V_2$$

This equation establishes the relationship between pressure and volume of the gas.

Experimental Verification of Boyle's law

The relationship between volume and pressure can be verified experimentally by the following series of experiments. Let us take some mass of a gas in a cylinder having a movable piston and observe the effect of increase of pressure on its volume. The phenomenon is represented in figure.5.1. When the pressure of 2 atmosphere (*atm*) is applied, the volume of the gas reads as 1 dm^3. When pressure is increased equivalent to 4 *atm*, the volume of the gas reduces to 0.5 dm^3. Again when pressure is increased three times i.e. 6 *atm*, the volume reduces to 0.33 dm^3. Similarly, when pressure is increased up to 8 *atm* on the piston, volume of the gas decreases to 0.25 dm^3.

Fig. 5.1 The decrease of volume with increase of pressure.

When we calculate the product of volume and pressure for this experiment, the product of all these experiments is constant i.e. 2 atm dm^3. It proves the Boyle's law

$$P_1V_1 = 2\,atm \times 1\,dm^3 = 2\,atm\,dm^3$$
$$P_2V_2 = 4\,atm \times 0.5\,dm^3 = 2\,atm\,dm^3$$
$$P_3V_3 = 6\,atm \times 0.33\,dm^3 = 2\,atm\,dm^3$$
$$P_4V_4 = 8\,atm \times 0.25\,dm^3 = 2\,atm\,dm^3$$

Test yourself 5.2

i. Is the Boyle's law applicable to liquids?
ii. Is the Boyle's law valid at very high temperature?
iii. What will happen if the pressure on a sample of gas is raised three times and its temperature is kept constant?

Do you know?

In which units blood pressure is measured?

Blood pressure is measured using a pressure gauge. It may be a mercury manometer or some other device. Blood pressure is reported by two values, such as 120/80, which is a normal blood pressure. The first measurement shows the maximum pressure when the heart is pumping. It is called ***systolic*** *pressure. When the heart is in resting position, pressure decreases and it is the second value called* ***diastolic***. *Both of these pressures are measured in torr units.* ***Hypertension*** *is because of high blood pressure due to tension and worries in daily life. The usual criterion for hypertension is a blood pressure greater than 140/90. Hypertension raises the level of stress on the heart and on the blood vessels. This stress increases the susceptibility of heart attacks and strokes.*

Example 5.1

A gas with volume 350 cm^3 has a pressure of 650 mm of Hg. If its pressure is reduced to 325 mm of Hg, calculate what will be its new volume?

Data

V_1 = 350 cm^3
P_1 = 650 *mm of Hg*
P_2 = 325 *mm of Hg*
V_2 = ?

Solution

By using the equation of Boyle's Law

$$P_1V_1 = P_2V_2$$

$$\text{or } V_2 = \frac{P_1V_1}{P_2}$$

By putting the values

$$V_2 = \frac{650 \times 350}{325}$$

$$= 700 \ cm^3$$

Thus volume of the gas is doubled by reducing its pressure to half.

Example 5.2

785 cm^3 of a gas was enclosed in a container under a pressure of 600 mm Hg. If volumes is reduced to 350 cm^3, what will be the pressure?

Data

V_1 = 785 cm^3
P_1 = 600 *mm of Hg*
V_2 = 350 cm^3
P_2 = ?

Solution

By using the Boyle's equation

$$P_1V_1 = P_2V_2$$

$$or \quad P_2 = \frac{P_1V_1}{V_2}$$

By putting the values

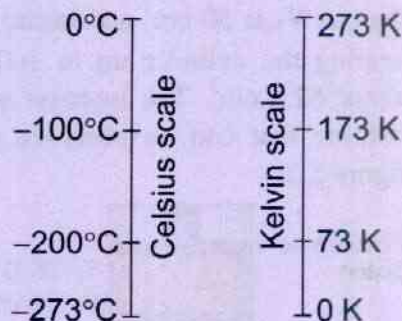
$$P_2 = \frac{785 \times 600}{350} = 1345.7 \ mm \ of \ Hg$$

$$or \quad P_2 = \frac{1345.7}{760} = 1.77 \ atm$$

Thus pressure is increased by decreasing volume.

Absolute Temperature Scale

Lord **Kelvin** introduced **absolute temperature scale** or **Kelvin scale**. This scale of temperature starts from 0 K or –273.15 °*C*, which is given the name of **absolute zero**. *It is the temperature at which an ideal gas would have zero volume.* As both scales have equal degree range, therefore, when 0 K equal to –273 °C then 273 K is equal to 0 °C as shown in the scales.

Conversion of Kelvin temperature to Celsius temperature and vice versa can be carried out as follows:

$$(T)\,K = (T)\,°C + 273$$
$$(T)\,°C = (T)\,K - 273$$

5.2.2 Charles's Law

The relationship between volume and temperature keeping the pressure constant was also studied. French scientist **J. Charles** in 1787 presented his law that states "*the volume of a given mass of a gas is directly proportional to the absolute temperature if the pressure is kept constant*". When pressure ***P*** is constant, the volume ***V*** of a given mass of a gas is proportional to absolute temperature ***T***. Mathematically it is represented as:

Volume ∝ temperature represented as $V \propto T$

or $V = kT$ or $\frac{V}{T} = k$

Where ***k*** is proportion... ...nt. If temperature of the gas is increased its volume also increases. When temperature is changed from T_1 to T_2 the volume changes from V_1 to V_2. The mathematical form of Charles' Law will be:

$$V_1 / T_1 = k \quad \text{and} \quad V_2 / T_2 = k$$

As both equations have same value of constant, therefore, their variables are also equal to each other

$$\therefore \quad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Experimental Verification of Charles's Law

Let us take a certain amount of gas enclosed in a cylinder having a movable piston. If the initial volume of the gas V_1 is 50 cm^3 and initial temperature T_1 is 25 °C, on heating the cylinder up to 100 °C, its new volume V_2 is about 62.5 cm^3. The increase in temperature, increases the volume that can be observed as elaborated below in the figure 5.2.

J. Charles (1746-1823) was a French inventor, scientist, mathematician and balloonist. He described in 1802, how gases tend to expand when heated.

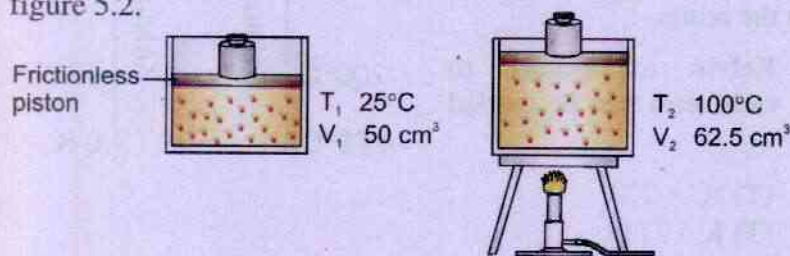



Fig. 5.2: Representation of increase of volume with the increase in temperature.

Remember

- Always convert temperature scale from °C to K scale while solving problems. K = 273 + °C

Example 5.3

A sample of oxygen gas has a volume of 250 cm^3 at –30 °C. If gas is allowed to expand up to 700 cm^3 at constant pressure, find out its final temperature.

Data

V_1 = 250 cm^3

T_1 = –30 °C = (–30 + 273) = 243 K

V_2 = 700 cm^3

T_2 = ?

Solution

By using the equation

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$or \quad T_2 = \frac{V_2 T_1}{V_1}$$

By putting the value in equation

$$T_2 = \frac{700 \times 243}{250} = 680.4\ K$$

Thus expansion is caused due to increasing temperature.

Example 5.4

A sample of hydrogen gas occupies a volume 160 cm^3 at 30 °C. If its temperature is raised to 100 °C, calculate what will be its volume if the pressure remains constant.

Data

V_1 = 160 cm^3

T_1 = 30 °C = 303 K (as 0 °C = 273 K)

T_2 = 100 °C = 373 K

V_2 = ?

Solution

By using the equation of Charles' Law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$or \quad V_2 = \frac{V_1 T_2}{T_1}$$

By putting the values in above equation

$$V_2 = \frac{160 \times 373}{303} = 196.9 \ cm^3$$

Thus volume of the gas has increased by raising the temperature.

Remember!

- Degree sign (∘) is used with Celsius scale not with Kelvin scale.

Test yourself 5.3

i. Which variables are kept constant in Charles's law?
ii. Why volume of a gas decreases with increase of pressure?
iii. What is absolute zero?
iv. Does Kelvin scale show a negative temperature?
v. When a gas is allowed to expand, what will be its effect on its temperature?
vi. Can you cool a gas by increasing its volume?

Do you know?

In which units' body temperature is measured?

Body temperature is measured in Fahrenheit scales. Normal body temperature is 98.6 °F, it is equivalent to 37 °C. This temperature is close to average normal atmospheric temperature. In winter atmospheric temperature falls lower than that of our body temperature.

According to principle of heat flow, heat flows out from our body and we feel cold. To control this outward flow of heat, we wear black and warm clothes. To maintain body temperature we use dry fruits, tea, coffee and meats etc.

Physical States of Matter and Role of Intermolecular Forces

As you know that matter exists in three physical states; gas, liquid and solid. In the gaseous state, the molecules are far apart from each other. Therefore, intermolecular forces are very weak in them. But in the liquid and solid states intermolecular forces play a very important role on their properties.

In the liquid state molecules are much closer to each other as compared to gases as shown in figure 5.3. As a result liquid molecules develop stronger intermolecular forces, which affect their physical properties like diffusion, evaporation, vapour pressure and boiling point. Compounds having stronger intermolecular forces have higher boiling points, as you will see in section 5.3.3.

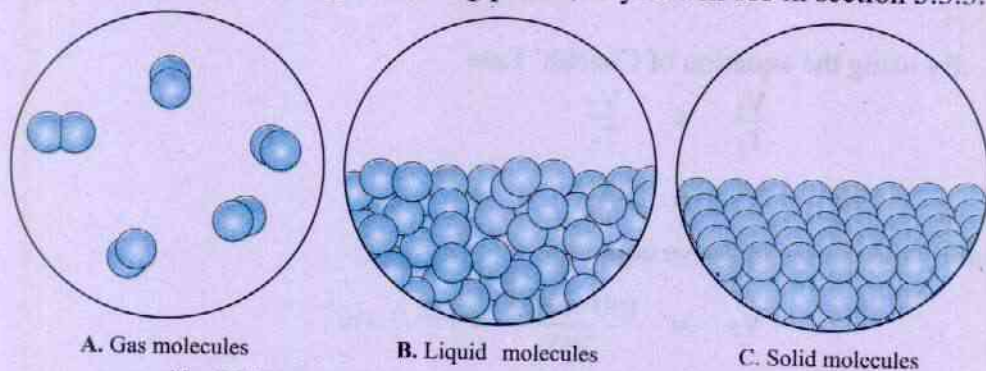



Fig 5.3 Three states of matter showing intermolecular forces.

The intermolecular forces become so dominant in solid state that the molecules look motionless. They arrange in a regular pattern therefore they are denser than molecules of liquids.

LIQUID STATE

Liquids have a definite volume but their shape is not definite. A liquid attains shape of the container in which it is put. A few typical properties of the liquids are discussed here.

5.3 TYPICAL PROPERTIES

5.3.1 Evaporation

The process of changing of a liquid into a gas phase is called ***evaporation***. It is reverse to condensation in which a gas changes into liquid. Evaporation is an endothermic process (heat is absorbed). Such as when one mole of water in liquid state is converted into vapour form, it requires 40.7 kJ of energy.

$$H_2O_{(l)} \longrightarrow H_2O_{(g)} \qquad \Delta H^\circ_{vap} = 40.7 kJ\ mol^{-1}$$

In the liquid state, molecules are in a continuous state of motion. They possess kinetic energy but all the molecules do not have same kinetic energy. Majority of the molecules have average kinetic energy and a few have more than average kinetic

energy. The molecules having more than average kinetic energy overcome the attractive forces among the molecules and escape from the surface. It is called as evaporation.

Evaporation is a continuous process taking place at all temperatures. The rate of evaporation is directly proportional to temperature. It increases with the increase in temperature because of increase in kinetic energy of the molecules.

Evaporation is a cooling process. When the high kinetic energy molecules vapourize, the temperature of remaining molecules falls down. To compensate this deficiency of energy, the molecules of liquid absorb energy from the surroundings. As a result the temperature of surroundings decreases and we feel cooling. For example, when we put a drop of alcohol on palm, the alcohol evaporates and we feel cooling effect.

Evaporation depends upon following **factors**:

i. ***Surface area***: Evaporation is a surface phenomenon. Greater is surface area, greater is evaporation and vice verse. For example, sometimes a saucer is used if tea is to be cooled quickly. This is because evaporation from the larger surface area of saucer is more than that from the smaller surface area of a tea cup.

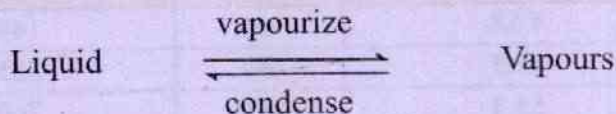
ii. ***Temperature***: At high temperature, rate of evaporation is high because at high temperature kinetic energy of the molecules increases so high that they overcome the intermolecular forces and evaporate rapidly. For example, water level in a container with hot water decreases earlier than that of a container with cold water. This is because the hot water evaporates earlier than the cold water.

iii. ***Intermolecular forces***: If intermolecular forces are stronger, molecules face difficulty in evaporation. For example, water has stronger intermolecular forces than alcohol, therefore, alcohol evaporates faster than water.

5.3.2 Vapour Pressure

*The pressure exerted by the vapours of a liquid at equilibrium with the liquid at a particular temperature is called **vapour pressure** of a liquid.*

The equilibrium is a state when rate of vapourization and rate of condensation is equal to each other but in opposite directions.

$$\text{Liquid} \underset{\text{condense}}{\overset{\text{vapourize}}{\rightleftharpoons}} \text{Vapours}$$

From the open surface of a liquid, molecules evaporate and mix up with the air but when we close a system, evaporated molecules start gathering over the liquid surface. Initially the vapours condense slowly to return to liquid. After sometime condensation process increases and a stage reaches when the rate of evaporation becomes equal to rate of condensation. At that stage the number of molecules

evaporating will be equal to the number of molecules coming back(condensing) to liquid. This state is called **dynamic equilibrium** as shown in figure 5.4.

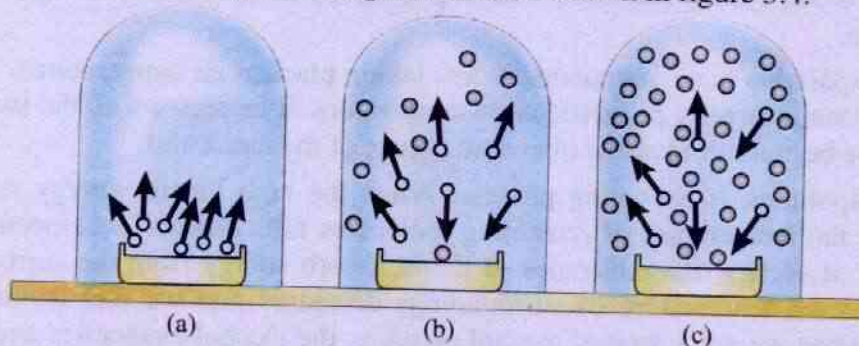



Fig. 5.4 A state of Dynamic Equilibrium between liquid and its vapours

Vapour pressure of a liquid depends upon the following factors.

i. ***Nature of liquid***: Vapour pressure depends upon the nature of liquid. Polar liquids have low vapour pressure than non-polar liquids at the same temperature. This is because of strong intermolecular forces between the polar molecules of liquids. For example, water has less vapour pressure than that of alcohol at same temperature.

ii. ***Size of molecules***: Small sized molecules can easily evaporate than big sized molecules hence, small sized molecular liquids exert more vapour pressure. For example, hexane (C_6H_{14}) has a small sized molecule as compared to decane ($C_{10}H_{22}$). Therefore C_6H_{14} evaporates rapidly and exerts vapour more pressure than $C_{10}H_{22}$.

iii. ***Temperature***: At high temperature, vapour pressure is higher than at low temperature. At elevated temperature, the kinetic energy of the molecules increases enough to enable them to vaporize and exert pressure.

For example, vapour pressure of water at different temperatures is given in the Table 5.1

Table 5.1 Relationship of Vapour Pressure of Water with Temperature

Temp °C	Vapour Pressure *mmHg*	Temp °C	Vapour Pressure *mmHg*
0	4.58	60	149.4
20	17.5	80	355.1
40	55.3	100	760.0

5.3.3 Boiling Point

When a liquid is heated, its molecules gain energy. The number of molecules which have more than average kinetic energy increases. More and more molecules become energetic enough to overcome the intermolecular forces. Due to this, rate of

evaporation increases that results in increase of vapour pressure until a stage reaches where the vapour pressure of a liquid becomes equal to atmospheric pressure. At this stage the liquid starts boiling. Hence **boiling point** *is defined as the temperature at which the vapour pressure of a liquid becomes equal to the atmospheric pressure or any external pressure.*

The figure 5.5 shows the increase of vapour pressure of diethyl ether, ethyl alcohol and water with the increase of temperature. At 0°C the vapour pressure of diethyl ether is 200 mm Hg, of ethyl alcohol 25 mm Hg while that of water is about 5 mm Hg. When they are heated, vapour pressure of diethyl ether increases rapidly and becomes equal to atmospheric pressure at 34.6°C, while vapour pressure of water increases slowly because intermolecular forces of water are stronger. The figure shows the vapour pressure increases very rapidly when the liquids are near to boiling point.

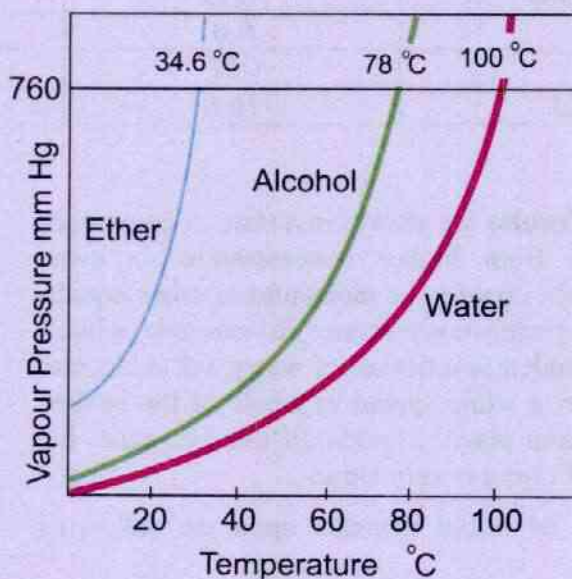



Fig. 5.5 Boiling point curves of Ether, Alcohol and Water.

The boiling point of the liquid depends upon the following factors.

i. ***Nature of liquid:*** The polar liquids have higher boiling points than that of non-polar liquids because polar liquids have strong intermolecular force. Boiling points of a few liquids are given in the table 5.2

ii. ***Intermolecular forces:*** Intermolecular forces play a very important role on the boiling point of liquids. Substances having stronger intermolecular forces have high boiling points, because such liquids attain a level of vapour pressure equal to external pressure at high temperature. It is given in figure 5.5.

iii. ***External pressure:*** Boiling points of a liquid depends upon external pressure. Boiling point of a liquid is controlled by external pressure in such a way, that

it can be increased by increasing external pressure and vice versa. This principle is used in the working of 'Pressure Cooker'.

5.3.4 Freezing Point

When liquids are cooled the vapour pressure of liquid decreases and a stage reaches when vapour pressure of a liquid state becomes equal to the vapour pressure of the solid state. At this temperature liquid and solid coexist in dynamic equilibrium and this is called the **freezing point** of a liquid. Boiling point and freezing point of a few liquids are given in the table 5.2

Table 5.2 Freezing and Boiling Points of Common Liquids

Sr. No	Liquid	Freezing Point °C	Boiling Point °C
1	Diethyl ether	–116	34.6
2	Ethyl alcohol	–115	78
3	Water	0.0	100
4	n-Octane	–57	126
5	Acetic acid	16.6	118

5.3.5 Diffusion

The liquid molecules are always in a state of continuous motion. They move from higher concentration to lower concentration. They mix up with the molecules of other liquids, so that they form a homogeneous mixture. For example, when a few drops of ink are added in a beaker of water, ink molecules move around and after a while spread in whole of the beaker. Thus diffusion has taken place. Liquids diffuse like gases but the rate of diffusion of liquid is very slow.

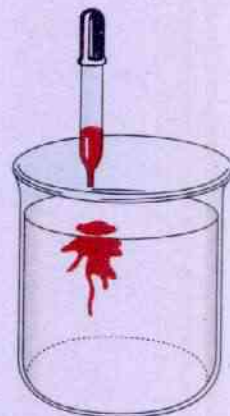
Fig. 5.6 Diffusion in liquids

The diffusion of liquid depends upon the following fac ors.

i. ***Intermolecular forces:*** Liquids having weak intermolecular forces diffuse faster than those having strong intermolecular forces.

ii. ***Size of molecules***: Big sized molecules diffuse slowly. For example, honey diffuses slowly in water than that of alcohol in water.

iii. ***Shapes of molecules***: Regular shaped molecules diffuse faster than irregular shaped molecules because they can easily slip over and move faster.

iv. ***Temperature:*** Diffusion increases by increasing temperature because at high temperature the intermolecular forces become weak due to high kinetic energy of the molecules.

5.3.6 Density

The density of liquid depends upon its mass per unit volume. Liquids are denser than gases because molecules of liquid are closely packed and the spaces between their molecules are negligible. As the liquid molecules have strong intermolecular forces hence they cannot expand freely and have a fixed volume. Like gases, they cannot occupy all the available volume of the container that is the reason why densities of liquids are high. For example: density of water is 1.0 $g\ cm^{-3}$ while that of air is 0.001 $g\ cm^{-3}$. That is the reason why drops of rain fall downward. The densities of liquids also vary. You can observe kerosene oil floats over water while honey settles down in the water.

Test yourself 5.4

- i. *Why does evaporation increase with the increase of temperature?*
- ii. *What do you mean by condensation?*
- iii. *Why is vapour pressure higher at high temperature?*
- iv. *Why is the boiling point of water higher than that of alcohol?*
- v. *What do you mean by dynamic equilibrium?*
- vi. *Why are the rates of diffusion in liquids slower than that of gases?*
- vii. *Why does rate of diffusion increase with increase of temperature?*
- viii. *Why are the liquids mobile?*

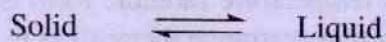
SOLID STATE

It is third state of matter which has definite shape and volume. In solid state the molecules are very close to one another and they are closely packed. The intermolecular forces are so strong that particles become almost motionless. Hence they cannot diffuse. Solid particles possess only vibrational motion.

5.4 TYPICAL PROPERTIES

Solids exhibit typical properties, a few of which are discussed here.

5.4.1 Melting point

The solid particles possess only vibrational kinetic energy. When solids are heated, their vibrational energies increase and particles vibrate at their mean position with a higher speed. If the heat is supplied continuously, a stage reaches at which the particles leave their fixed positions and then become mobile. At this temperature solid **melts**. The *temperature at which the solid starts melting and coexists in dynamic equilibrium with liquid state is called* ***melting point***. The ionic and covalent solids make network structure to form macromolecules. So all such solids have very high melting points.

$$\text{Solid} \rightleftharpoons \text{Liquid}$$

5.4.2 Rigidity

The particles of solids are not mobile. They have fixed positions. Therefore, solids are rigid in their structure.

5.4.3 Density

Solids are denser than liquids and gases because solid particles are closely packed and do not have empty spaces between their particles. Therefore, they have the highest densities among the three states of matter. For example, density of aluminium is 2.70 $g\ cm^{-3}$, iron is 7.86 $g\ cm^{-3}$ and gold is 19.3 $g\ cm^{-3}$.

5.5 Types of Solids

According to their general appearance solids can be classified into two types: amorphous solids and crystalline solids.

5.5.1 Amorphous Solids

Amorphous means shapeless. *Solids in which the particles are not regularly arranged or their regular shapes are destroyed, are called **amorphous solids***. They do not have sharp melting points. Plastic, rubber and even glass are amorphous solids as they do not have any sharp melting points.

5.5.2 Crystalline Solids

*Solids in which particles are arranged in a definite three-dimensional pattern are called **crystalline solids***. They have definite surfaces or faces. Each face has definite angle with the other. They have sharp melting points. Examples of crystalline solids are diamond, sodium chloride etc.

5.6 Allotropy

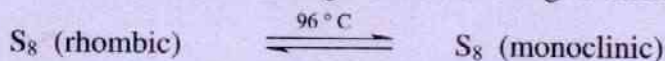
The existence of an element in more than one forms in same physical state is called **allotropy**. Allotropy is due to:

i. The existence of two or more kinds of molecules of an element each having different number of atoms such as allotropes of oxygen are oxygen (O_2) and ozone (O_3)

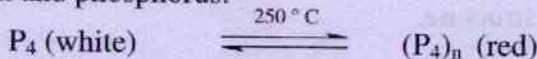
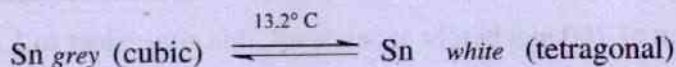
ii. Different arrangement of two or more atoms or molecules in a crystal of the element. Such as, sulphur shows allotropy due to different arrangement of molecules (S_8) in the crystals.

They always show different physical properties but have same chemical properties.

Allotropes of solids have different arrangement of atoms in space at a given temperat e. The arrangement of atoms also change with the change of temperature and new .llotropic form is produced. *The temperature at which one allotrope changes into another is called **transition temperature***. For example, transition temperature of sulphur is 96 °C. Below this temperature rhombic form is stable. If rhombic form is heated above 96 °C, its molecules rearrange themselves to give monoclinic form.

$$S_8\ (\text{rhombic}) \xrightleftharpoons{96\,^\circ C} S_8\ (\text{monoclinic})$$

Other examples are tin and phosphorus.

$$P_4\ (\text{white}) \xrightleftharpoons{250\,^\circ C} (P_4)_n\ (\text{red})$$

$$\text{Sn } grey \text{ (cubic)} \xrightleftharpoons{13.2°\,C} \text{Sn } white \text{ (tetragonal)}$$

White phosphorus is very reactive, poisonous and waxy solid. It exists as tetra-atomic molecules. While red phosphorous is less reactive, non-poisonous and a brittle powder.

Test yourself 5.5

i. Which form of sulphur exists at room temperature?
ii. Why is white tin available at room temperature?
iii. Why is the melting point of a solid considered its 'identification' characteristic?
iv. Why amorphous solids do not have sharp melting points while crystalline solids do have?
v. Which is lighter one aluminium or gold?
vi. Write the molecular formula of a sulphur molecule?
vii. Which allotropic form of carbon is stable at room temperature (25 °C)?
viii. State whether allotropy is shown by elements or compounds or both?

Curing with salt to preserve meat

Table salt is the most important ingredient for curing meat and is used in large quantities. Salt kills and inhibits the growth of putrifying bacteria by drawing water out of the meat. Concentrations of salt up to 20% are required to kill most species of unwanted bacteria. Once properly salted, the meat contains enough salt to prevent the growth of many undesirable microbes.

CHANGE OF INSTRUMENTATION AS THE SCIENCE PROGRESSES

There are many aspects to be considered about the functioning of instruments. Scientific observation is determined by the human sensory system. It generally relies on instruments that serve as mediators between the world and the senses. Thus, instruments can be considered as a reinforcement of the senses. They provide a great capacity for increasing the power of observation and making induction processes easier. Furthermore, scientific instruments constitute a major factor in checking, refuting or changing previously established theories.

KEY POINTS

- Gases diffuse very rapidly. Diffusion is mixing up of a gas throughout a space or other gases.
- Effusion is escaping of a gas molecule through a fine hole into an evacuated space.
- Gases exert pressure. The SI unit of pressure is Nm^{-2} which is also called Pascal .
- Standard atmospheric pressure is the pressure exerted by a mercury

column of 760 *mm* height at sea level, it is equivalent to 1 *atmosphere*.

- Gases are highly mobile and they can be compressed.
- Gases are 1000 times lighter than liquids or solids hence their density is measured in gdm^{-3}.
- Boyle's law states that volume of a given mass of a gas is inversely proportional to the pressure at constant temperature.
- Charles' Law states that volume of a given mass of a gas is directly proportional to the absolute temperature at a constant pressure.
- Absolute zero is the temperature at which an ideal gas would have zero volume , it is –273.15 °C.
- The conversion of a liquid into vapours at all temperatures is called evaporation. It is a cooling process.
- Evaporation depends upon surface area, temperature and intermolecular forces.
- Vapour pressure of a liquid is defined as the pressure exerted by the vapours when liquid and vapour states are in dynamic equilibrium with each other.
- Boiling point is the temperature at which the vapour pressure of a liquid becomes equal to the atmospheric pressure or any external pressure.
- Boiling point depends upon the nature of liquid, intermolecular forces and external pressure.
- Freezing point of a liquid is that temperature at which vapour pressure of liquid phase is equal to the vapour pressure of the solid phase. At this temperature liquid and solid coexist in dynamic equilibrium with one another.
- Melting point of solid is the temperature at which solid when heated melts and coexist in dynamic equilibrium with liquid.
- Solids are rigid and denser than liquids.
- Solids are classified as amorphous and crystalline .
- Amorphous solids are shapeless and do not have sharp melting point.
- Crystalline solids have definite three dimensional pattern of arrangement of particles .They have sharp melting points.
- The existence of a solid in different physical forms is called allotropy.

EXERCISE

Multiple Choice Questions

Put a (✓) on the correct answer

1. **How many times liquids are denser than gases?**
   (a) 100 times (b) 1000 times
   (c) 10,000 times (d) 100,000 times
2. **Gases are the lightest form of matter and their densities are expressed in terms of:**
   (a) mg cm^{-3} (b) g cm^{-3} (c) g dm^{-3} (d) kg dm^{-3}
3. **At freezing point which one of the following coexists in dynamic equilibrium:**
   (a) gas and solid (b) liquid and gas
   (c) liquid and solid (d) all of these
4. **Solid particles possess which one of the following motions?**
   (a) rotational motions (b) vibrational motions
   (c) translational motions (d) both translational and vibrational motions
5. **Which one of the following is not amorphous?**
   (a) rubber (b) plastic (c) glass (d) glucose
6. **One atmospheric pressure is equal to how many Pascals:**
   (a) 101325 (b) 10325 (c) 106075 (d) 10523
7. **In the evaporation process, liquid molecules which leave the surface of the liquid have:**
   (a) very low energy (b) moderate energy
   (c) very high energy (d) none of these
8. **Which one of the following gas diffuses fastest?**
   (a) hydrogen (b) helium
   (c) fluorine (d) chlorine
9. **Which one of the following does not affect the boiling point?**
   (a) intermolecular forces (b) external pressure
   (c) nature of liquid (d) initial temperature of liquid
10. **Density of a gas increases, when its:**
    (a) temperature is increased (b) pressure is increased
    (c) volume is kept constant (d) none of these
11. **The vapour pressure of a liquid increases with the:**
    (a) increase of pressure
    (b) increase of temperature
    (c) increase of intermolecular forces
    (d) increase of polarity of molecules

Short Answer Questions

1. What is diffusion, explain with an example?
2. Define standard atmospheric pressure. What are its units? How it is related to Pascal?
3. Why are the densities of gases lower than that of liquids?
4. What do you mean by evaporation how it is affected by surface area.
5. Define the term allotropy with examples.
6. In which form sulphur exists at 100 °C.
7. What is the relationship between evaporation and boiling point of a liquid?

Long Answer Questions

1. Define Boyle's law and verify it with an example.
2. Define and explain Charles' law of gases.
3. What is vapour pressure and how it is affected by intermolecular forces.
4. Define boiling point and also explain, how it is affected by different factors.
5. Describe the phenomenon of diffusion in liquids along with factors which influence it.
6. Differentiate between crystalline and amorphous solids.

Numerical

1. Convert the following units:
   (a) 850 mm Hg to atm (b) 205000 Pa to atm
   (c) 560 torr to cm Hg (d) 1.25 atm to Pa
2. Convert the following units:
   (a) 750 °C to K (b) 150 °C to K
   (c) 100 K to °C (d) 172 K to °C.
3. A gas at pressure 912 mm of Hg has volume 450cm^3. What will be its volume at 0.4 atm.
4. A gas occupies a volume of 800 cm^3 at 1 atm, when it is allowed to expand up to 1200 cm^3 what will be its pressure in mm of Hg.
5. It is desired to increase the volume of a fixed amount of gas from 87.5 to 118 cm^3 while holding the pressure constant. What would be the final temperature if the initial temperature is 23 °C.

6. A sample of gas is cooled at constant pressure from 30 °C to 10 °C. Comment:
   a. Will the volume of the gas decrease to one third of its original volume?
   b. If not, then by what ratio will the volume decrease?

7. A balloon that contains 1.6 dm^3 of air at standard temperature (0 °C) and (1atm) pressure is taken under water to a depth at which its pressure increases to 3.0 atm. Suppose that temperature remain unchanged, what would be the new volume of the balloon. Does it contract or expand?

8. A sample of neon gas occupies a volume of 75.0 cm^3 at very low pressure of 0.4 atm. Assuming temperature remain constant what would be the volume at 1.0 atm. pressure?

9. A gas occupies a volume of 35.0 dm^3 at 17 °C. If the gas temperature rises to 34°C at constant pressure, would you expect the volume to double? If not calculate the new volume.

10. The largest moon of Saturn, is Titan. It has atmospheric pressure of 1.6×10^5 Pa. What is the atmospheric pressure in atm? Is it higher than earth's atmospheric pressure?

Solutions

Major Concepts

6.1 Solution, aqueous solution, solute and solvent

6.2 Saturated, unsaturated, supersaturated solutions and dilution of solution

6.3 Types of solutions

6.4 Concentration units

6.5 Comparison of solutions, suspensions and colloids

Time allocation

Teaching periods	16
Assessment periods	02
Weightage	14%

Students Learning Outcomes

Students will be able to:

- Define the terms: solution, aqueous solution, solute and solvent and give an example of each.
- Explain the difference between saturated, unsaturated and supersaturated solutions.
- Explain the formation of solutions (mixing gases into gases, gases into liquids, gases into solids) and give an example of each.
- Explain the formation of solutions (mixing liquids into gases, liquids into liquids, liquids into solids) and give an example of each.
- Explain the formation of solutions (mixing solids into gases, solids into liquids, solids into solids) and give an example of each.
- Explain what is meant by the concentration of a solution.
- Define molarity.
- Define percentage solution.
- Solve problems involving the molarity of solution.
- Describe how to prepare dilute solutions from concentrated solutions of known molarity.
- Convert between the molarity of a solution and its concentration in g/dm^3.
- Use the rule that "like dissolves like" to predict the solubility of one substance in another.

Introduction

Solutions are homogeneous mixtures of two or more components. Generally, solutions are found in three physical states depending upon the physical state of the solvent, e.g. alloy is a solid solution; sea water is a liquid solution and air is a gaseous solution. There are nine types of solutions ranging from gas-gas e.g air we breathe to solid-solid solutions e.g dental amalgam for filling of tooth. Liquid solutions are the most common solutions because of the most common solvent water. Therefore, there is a wide variety of liquid solutions ranging from a drop of rain to oceans. Sea water is a resource of 92 naturally occurring elements.

6.1 SOLUTION

A solution *is a homogeneous mixture of two or more substances.* The boundaries of the components can't be distiguished i.e. a solution exist as one phase. For example, the air we breathe is a solution of several gases, brass is a solid solution of Zn and Cu. Sugar dissolved in water is an example of liquid solution.

The simplest way to distinguish between a solution and a pure liquid is evaporation. The liquid which evaporates completely, leaving no residue, is a pure compound, while a liquid which leaves behind a residue on evaporation is solution. An alloy like brass or bronze is also a homogeneous mixture. Although it cannot be separated by physical means, yet it is considered a mixture as:

i. It shows the properties of its components and

ii. It has a variable composition.

6.1.1 Aqueous Solutions

The solution which is formed by dissolving a substance in water is called an **aqueous solution**. In aqueous solutions water is always present in greater amount and termed as *solvent*. For example, sugar in water and table salt in water. Aqueous solutions are mostly used in the laboratories. Water is called a universal solvent because it dissolves majority of compounds present in earth's crust.

6.1.2 Solute

The component of solution which is present in smaller quantity is called **solute**. A solute is dissolved in a solvent to make a solution. For example, salt solution is made by dissolving salt in water. So in salt solution, salt is the solute and water is solvent. More than one solutes may be present in a solution. For example, in soft drinks, water is a solvent .while other substances like sugar, salts and CO_2 are solutes.

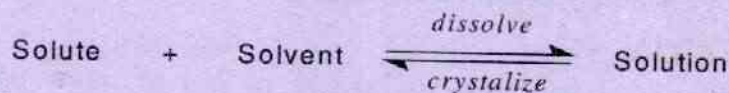
6.1.3 Solvent

The component of a solution which is present in larger quantity is called **solvent**. Solvent always dissolves solutes. In a solution if more than two substances are present, one substance acts as solvent and others behave as solutes. For example,

as referred above in soft drinks, water is solvent while other substances like sugar, salts and CO_2 are solutes.

6.2 SATURATED SOLUTION

When a small amount of solute is added in a solvent, solute dissolves very easily in the solvent. If the addition of solute is kept on, a stage is reached when solvent cannot dissolve any more solute. At this stage, further added solute remains undissolved and it settles down at the bottom of the container.

$$\text{Solute} + \text{Solvent} \underset{crystalize}{\overset{dissolve}{\rightleftharpoons}} \text{Solution}$$

A solution containing maximum amount of solute at a given temperature is called **saturated solution**. On the particle level, a saturated solution is the one, in which undissolved solute is in equilibrium with dissolved solute.

$$\text{Solute (crystallized)} \rightleftharpoons \text{Solute (dissolved)}$$

At this stage dynamic equilibrium is established. Although dissolution and crystallization continues at a given temperature, but the net amount of dissolved solute remains constant.

6.2.1 Unsaturated Solution

A solution which contains lesser amount of solute than that which is required to saturate it at a given temperature, is called **unsaturated solution**. Such solutions have the capacity to dissolve more solute to become a saturated solution.

6.2.2 Supersaturated Solution

When saturated solutions are heated, they develop further capacity to dissolve more solute. Such solutions contain greater amount of solute than is required to form a saturated solution and they become more concentrated. The *solution that is more concentrated than a saturated solution is known as* **supersaturated solution**. Super-saturated solutions are not stable. Therefore, an easy way to get a supersaturated solution is to prepare a saturated solution at high temperature. It is then cooled to a temperature where excess solute crystallizes out and leaves behind a saturated solution. For example, a saturated solution of sodium thiosulphate ($Na_2S_2O_3$) in water at 20 °C has 20.9 g of salt per 100 cm^3 of water. Less than this amount of salt per 100 cm^3 of water at 20 °C will be an unsaturated solution. A solution having more amount than 20.9 g of salt per 100 cm^3 of water at 20 °C will be a supersaturated solutuion.

6.2.3 Dilution of Solution

The solutions are classified as dilute or concentrated on the basis of relative amount of solute present in them. **Dilute solutions** *are those which contain relatively small amount of dissolved solute in the solution.* **Concentrated solutions** *are those which contain relatively large amount of dissolved solute in the solution.* For example, brine is a concentrated solution of common salt in water. These terms describe the concentration of the solution. Addition of more solvent will dilute the solution and its concentration decreases.

The preparation of dilute solutions from concentrated solutions has been explained in Section 6.4.3.1.

6.3 TYPES OF SOLUTION

Each solution consists of two components, solute and solvent. The solute as well as solvent may exist as gas, liquid or solid. So, depending upon the nature of solute and solvent different types of solutions may form, which are given in table 6.1.

Table 6.1 Different Types of Solutions with Examples

Sr. No	Solute	Solvent	Example of Solution
1	Gas	Gas	Air, mixture of H_2 and He in weather balloons, mixture of N_2 and O_2 in cylinders for respiration.
2	Gas	Liquid	Oxygen in water, carbon dioxide in water.
3	Gas	Solid	Hydrogen adsorbed on palladium.
4	Liquid	Gas	Mist, fog, liquid air pollutants.
5	Liquid	Liquid	Alcohol in water, benzene in toluene.
6	Liquid	Solid	Butter, cheese.
7	Solid	Gas	Dust particles or smoke in air.
8	Solid	Liquid	Sugar in water.
9	Solid	Solid	Metal alloys(brass, bronze), opals.

Test yourself 6.1

i. *Why is a solution considered mixture?*
ii. *Distinguish between the following pairs as compound or solution:*
   (a) water and salt solution (b)vinegar and benzene
   (c) carbonated drinks and acetone
iii. *What is the major difference between a solution and a mixture?*
iv. *Why are the alloys considered solutions?*
v. *Dead sea is so rich with salts that it forms crystals when temperature lowers in the winter. Can you comment why is it named as "Dead Sea"?*

6.4 CONCENTRATION UNITS

Concentration is the proportion of a solute in a solution. It is also *a ratio of the amount of solute* to the amount of solution or ratio of amount of solute to the amount of the solvent. Please keep in mind that concentration does not depend upon the total volume or total amount of the solution. For example, a sample taken from the bulk solution will have the same concentration. There are various types of units used to express concentration of solutions. A few of these units are discussed here.

6.4.1 Percentage

Percentage unit of concentration refers to *the percentage of solute present in a solution*. The percentage of solute can be expressed by mass or by volume.

It can be expressed in terms of percentage composition by four different ways.

6.4.1.1 Percentage - mass/mass (%m/m)

It is the number of grams of solute in 100 grams of solution. For example, 10% m/m sugar solution means that 10 *g* of sugar is dissolved in 90 *g* of water to make 100 *g* of solution. Calculation of this ratio is carried out by using the following formula:

$$\%\ \text{mass/mass} = \frac{\textit{mass of solute}(g)}{\textit{mass of solute}(g) + \textit{mass of solvent}(g)} \times 100$$

$$= \frac{\textit{mass of solute}(g)}{\textit{mass of solution}(g)} \times 100$$

6.4.1.2 Percentage - mass/volume (%m/v)

It is the number of grams of solute dissolved in 100 cm^3 (parts by volume) of the solution. For example, 10 % **m/v** sugar solution contains 10 *g* of sugar in 100 cm^3 of the solution. The exact volume of solvent is not mentioned or it is not known.

$$\%\ \text{m/v} = \frac{\textit{Mass of solute (g)}}{\textit{Volume of solution } (cm^3)} \times 100$$

6.4.1.3 Percentage - volume/mass (%v/m)

It is the volume in cm^3 of a solute dissolved in 100 g of the solution. For example, 10 % **v/m** alcohol solution in water means 10 cm^3 of alcohol is dissolved in (unknown) volume of water so that the total mass of the solution is 100 *g*. In such solutions the mass of solution is under consideration, total volume of the solution is not considered.

$$\% \text{ v/m} = \frac{\textit{Volume of solute } (cm^3)}{\textit{Mass of solution } (g)} \times 100$$

6.4.1.4 Percentage - volume /volume (% v/v)

It is the volume in cm^3 of a solute dissolved per 100 cm^3 of the solution. For example, 30 percent alcohol solution means 30 cm^3 of alcohol dissolved in sufficient amount of water, so that the total volume of the solution becomes 100 cm^3.

$$\% \text{ volume/volume} = \frac{\text{Volume of solute (cm}^3)}{\text{Volume of solution (cm}^3)} \times 100$$

Example 6.1

If we add 5 cm^3 of acetone in water to prepare 90 cm^3 of aqueous solution. Calculate the concentration(v/v) of this solution.

Solution

Using the relationship

$$\% \text{ volume/volume} = \frac{\text{Volume of the solute}}{\text{Volume of the solution}} \times 100$$

$$= \frac{5}{90} \times 100 = 5.5$$

Thus concentration of solution is 5.5 percent by volume.

6.4.2 Molarity

It is a concentration unit defined as *number of moles of solute dissolved in one dm^3 of the solution.* It is represented by **M**. Molarity is the unit mostly used in chemistry and allied sciences. The formula used for the preparation of molar solution is as follows:

$$\text{Molarity } (M) = \frac{\dfrac{\textit{Mass of solute}(g)}{\textit{Molar mass of solute}(g mol^{-1})}}{\textit{Volume of solution}(dm^3)} = \frac{\text{No. of moles of solute}}{\text{Volume of solution (dm}^3)}$$

$$\textit{or} \quad \text{Molarity } (M) = \frac{\textit{Mass of solute}(g)}{(\textit{Molar mass of solute}(g mol^{-1})) \times (\textit{Volume of solution}(dm^3))}$$

6.4.2.1 Preparation of Molar Solution

One **Molar solution** is prepared by dissolving 1 *mole* (molar mass) of the solute in sufficient amount of water to make the total volume of the solution up to $1 dm^3$ in a measuring flask. For example, $1M$ solution of NaOH is prepared by dissolving 40 *g* of NaOH in sufficient water to make the total volume $1 dm^3$.

As amount of solute is increased, its concentration or molarity also increases. 2.0 M solution is more concentrated than 1.0 M solution.

Test yourself 6.2

i. Does the percentage calculations require the chemical formula of the solute?
ii. Why is the formula of solute necessary for calculation of the molarity of the solution?
iii. You are asked to prepare 15 percent (m/m) solution of common salt. How much amount of water will be required to prepare this solution?
iv. How much water should be mixed with 18 cm^3 of alcohol so as to obtain 18 % (v/v) alcohol solution?
v. Calculate the concentration % (m/m) of a solution which contains 2.5 g of salt dissolved in 50 g of water.
vi. Which one of the following solutions is more concentrated: one molar or three molar

6.4.3 Problems involving the molarity of a solution

The following solved examples will help you to understand how molar solutions are prepared.

Example 6.2

Calculate the molarity of a solution which is prepared by dissolving 28.4 g of Na_2SO_4 in 400 cm^3 of solution.

Solution

Conversion mass of solute into *moles*

$$\text{No of moles } Na_2SO_4 = \frac{mass\ dissolved\ (g)}{molar\ mass\ (gmol^{-1})}$$

$$= \frac{28.4\ g}{142\ gmol^{-1}} = 0.2\ mol$$

$$\text{Conversion of volume into } dm^3 = \frac{400\ cm^3}{1000\ cm^3} \times 1\ dm^3 = 0.4\ dm^3$$

$$\text{Molarity} = \frac{no.of\ moles}{volume\ of\ solution\ (dm^3)}$$

$$= \frac{0.2}{0.4} = 0.5\ \text{mol dm}^{-3}$$

Example 6.3

How much NaOH is required to prepare its 500 cm^3 of 0.4 M solution.

Solution

Molar mass of NaOH $= 40\ gmol^{-1}$

Volume in dm^3 $= \dfrac{500\ cm^3}{1000\ cm^3} \times 1\ dm^{-3}$

$= 0.5\ dm^3$

Putting the values in formula:

$$\text{Molarity} = \frac{mass\ of\ solute\ (g)}{molar\ mass\ (gmol^{-1}) \times volume\ of\ the\ solution\ (dm^3)}$$

Mass of solute = Molarity × molar mass × volume

$= 0.4 \times 40 \times 0.5$

$= 8\ g$

6.4.3.1 Dilution of Solution

Dilute molar solution is prepared from a concentrated solution of known molarity as explained below:

Suppose we want to prepare 100 cm^3 of 0.01 M solution from given 0.1 M solution of potassium permanganate. First 0.1 M solution is prepared by dissolving 15.8 g of potassium permanganate in 1 dm^3 of solution. Then 0.01 M solution is prepared by the dilution according to following calculations:

Fig. 6.1 Dilution of a solution.

Concentrated solution ***Dilute solution***

$M_1V_1 = M_2V_2$

Where $M_1 = 0.1\ M$

$V_1 = ?$

and

$V_2 = 100\ cm^3$

$M_2 = 0.01\ M$

Putting the values in above equation we get:

Concentrated solution ***Dilute solution***

$V_1 \times 0.1 = 0.01 \times 100$

$V_1 = \dfrac{0.01 \times 100}{0.1}$

$= 10\ cm^3$

Concentrated solution of $KMnO_4$ has dense purple colour. Take 10 cm^3 of this solution with the help of a graduated pipette and put in a measuring flask of 100 cm^3. Add water upto the mark present at the neck of the flask. Now it is 0.01 molar solution of $KMnO_4$.

Example 6.4

10 cm^3 of 0.01 molar $KMnO_4$ solution has been diluted to 100 cm^3. Find out the molarity of this solution.

Solution:

Data

$M_1 = 0.01\ M \qquad M_2 = ?$

$V_1 = 10\ cm^3 \qquad V_2 = 100\ cm^3$

Using following formula, volume required can be calculated

$$M_1V_1 = M_2V_2$$

$$or \quad M_2 = \frac{M_1V_1}{V_2}$$

By putting these values we get molarity:

$$M_2 = \frac{0.01 \times 10}{100} = 0.001\ M$$

6.5 SOLUBILITY

Solubility is defined as *the number of grams of the solute dissolved in 100 g of a solvent to prepare a saturated solution at a particular temperature*. The concentration of a saturated solution is referred to as solubility of the solute in a given solvent.

Following are the factors which affect the solubility of solutes:

1. The general principle of solubility is, like dissolves like.
   i. The ionic and polar substances are soluble in polar solvents. Ionic solids and polar covalent compounds are soluble in water e.g., KCl, Na_2CO_3, $CuSO_4$, sugar, and alcohol are all soluble in water.
   ii.. Non-polar substances are not soluble in polar solvents. Non-polar covalent compounds are not soluble in water such as ether, benzene, and petrol are insoluble in water.
   iii. Non-polar covalent substances are soluble in non-polar solvents (mostly organic solvents). Grease, paints, naphthalene are soluble in ether or carbon tetrachloride etc.
2. Solute-solvent interaction.
3. Temperature.

6.5.1 Solubility and Solute-solvent interaction

The solute-solvent interaction can be explained in terms of creation of attractive forces between the particles of solute and those of solvent. To dissolve one substance (solute) in another substance (solvent) following three events must occur :

i. Solute particles must separate from each other

ii. Solvent particles must separate to provide space for solute particles.

iii. Solute and solvent particles must attract and mix up.

Solution formation depends upon the relative strength of attractive forces between solute-solute, solvent-solvent and solute-solvent. Generally solutes are solids. Ionic solids are arranged in such a regular pattern that the inter-ionic forces are at a maximum. If the new forces between solute and solvent particles overcome the solute-solute attractive forces, then solute dissolves and makes a solution. If forces between solute particles are strong enough than solute-solvent forces, solute remains insoluble and solution is not formed. Figure 6.2 shows dissolution process by the interaction of solvent molecules with the solute ions. The solvent molecules first pull apart the solute ions and then surround them. In this way solute dissolves and solution forms.

Fig. 6.2 Interaction of solute and solvent to form solution.

For example, when NaCl is added in water it dissolves readily because the attractive interaction between the ions of NaCl and polar molecules of water are strong enough to overcome the Na^+ attractive forces between Na^+ and Cl^- ions in solid NaCl crystal. In this process the positive end of the water dipole is oriented towards the Cl^- ions and the negative end of water dipole is oriented towards the Na^+ ions. These ion-dipole attractions between Na^+ ions and water molecules, Cl^- ions and water molecules are so strong that they pull these ions from their positions in the crystal and thus NaCl dissolves. It is shown in the figure 6.2.

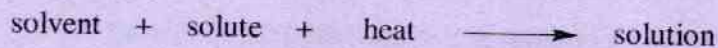
6.5.2 Effect of Temperature on solubility

Temperature has major effect on the solubility of most of the substances. Generally it seems that solubility increases with the increase of temperature, but it is not always true. When a solution is formed by adding a salt in solvent there are

different possibilities with reference to effect of temperature on solubility as shown in the figure 6.3. These possibilities are discussed here.

i. Heat is absorbed

When salts like KNO_3, $NaNO_3$ and KCl are added in water, the test tube becomes cold. It means during dissolution of these salts heat is absorbed. Such dissolving process is called '*endothermic*'.

solvent + solute + heat ⟶ solution

Solubility usually increases with the increase in temperature for such solutes. It means that heat is required to break the attractive forces between the ions of solute. This requirement is fulfilled by the surrounding molecules. As a result, their temperature falls down and test tube becomes cold.

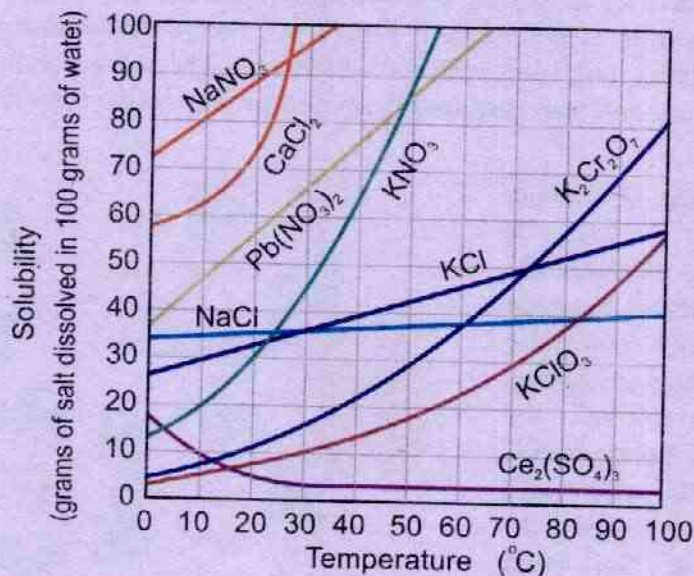
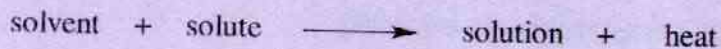



Fig. 6.3 Effect of temperature on solubility of different salts in water.

ii. Heat is given out

On the other hand when salts like Li_2SO_4 and $Ce_2(SO_4)_3$ are dissolved in water, the test tube becomes warm. i.e. heat is released during this dissolution.

solvent + solute ⟶ solution + heat

In such cases, the solubility of salt decreases with the increase of temperature. In such cases attractive forces among the solute particles are weaker and solute-solvent interactions are stronger. As a result, there is release of energy.

iii. No change in heat

In some cases during a dissolution process neither the heat is absorbed nor released. When salt like NaCl is added in water, the solution temperature remains almost the same. In such case temperature has a minimum effect on solubility .Figure 6.3 shows the trend of solubilities of different salts with the increase in temperature.

Test yourself 6.3

- *i. What will happen if the solute-solute forces are stronger than those of solute-solvent forces?*
- *ii. When solute-solute forces are weaker than those of solute-solvent forces? Will solution form?*
- *iii. Why is iodine soluble in CCl_4 and not in water.*
- *iv. Why test tube becomes cold when KNO_3 is dissolved in water*

6.6 COMPARISON OF SOLUTION, SUSPENSION AND COLLOID

6.6.1 Solution

Solutions are the homogeneous mixtures of two or more than two components. Each component is mixed in such a way that their individual identity is not visible. The simplest example is that of a drop of ink mixed in water. This is an example of true solution.

6.6.2 Colloid

These are solutions in which the solute particles are larger than those present in the true solutions but not large enough to be seen by naked eye. The particles in such system dissolve and do not settle down for a long time. *But particles of colloids are big enough to scatter the beam of light.* It is called ***Tyndall effect***. We can see the path of scattered light beam inside the colloidal solution. Tyndall effect is the main characteristic which distinguishes colloids from solutions. Hence, these solutions are called false solutions or colloidal solutions. Examples are starch, albumin, soap solutions, blood, milk, ink, jelly and toothpaste, etc.

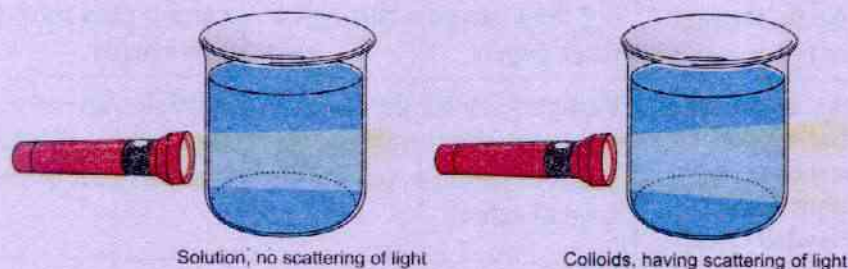



Fig. 6.4 Tyndall effect by colloids.

6.6.3 Suspension

Suspensions are a heterogeneous mixture of undissolved particles in a given medium. Particles are big enough to be seen with naked eyes. Examples are chalk in water(milky suspension), paints and milk of magnesia (suspension of magnesium oxide in water).

For better understanding of true solutions, false solution and suspension, a comparison of their characteristics is given in table 6.2.

Table 6.2 Comparison of the Characteristics of Solution, Colloidal and Suspension

Solution	Colloid	Suspension
The particles exist in their simplest form i.e. as molecules or ions. Their diameter is 10^{-8} cm.	The particles are large consisting of many atoms, ions or molecules.	The particles are of largest size. They are larger than 10^{-5} cm in diameter.
Particles dissolve uniformly throughout and form a homogeneous mixture.	A colloid appears to be a homogeneous but actually it is a heterogeneous mixture. Hence, they are not true solution. Particles do not settle down for a long time, therefore, colloids are quite stable	Particles remain undissolved and form a heterogeneous mixture. Particles settle down after sometime
Particles are so small that they can't be seen with naked eye.	Particles are large but can't be seen with naked eye.	Particles are big enough to be seen with naked eye.
Solute particles can pass easily through a filter paper.	Although particles are big but they can pass through a filter paper.	Solute particles cannot pass through filter paper.
Particles are so small that they cannot scatter the rays of light, thus do not show tyndall effect.	Particles scatter the path of light rays thus emitting the beam of light i.e. exhibit the tyndall effect.	Particles are so big that light is blocked and difficult to pass.

Test yourself 6.4

i. *What is difference between colloid and suspension?*
ii. *Can colloids be separated by filtration, if not why?*
iii. *Why are the colloids quite stable?*
iv. *Why does the colloid show tyndall effect?*
v. *What is tyndall effect and on what factors it depends?*
vi. *Identify as colloids or suspensions from the following: Paints, milk, milk of magnesia, soap solution.*
vii. *How can you justify that milk is a colloid.*

RELATIONSHIP OF SOLUTIONS TO DIFFERENT PRODUCTS IN THE COMMUNITY

Our body is made up of tissues, which are all composed of water based chemicals. The water becomes the best solvent in our body. We need an adequate supply of chemicals in the form of food, vitamins, hormones, and enzymes. For taking care of our health we need medicines. We find that chemicals and chemistry penetrate into every aspect of our life. Paper, sugar, starch, vegetable oils, ghee, essential oils, tannery, soap, cosmetics, rubber, dyes, plastics, petroleum, infact, there is almost nothing that we use in our daily life that is not a chemical. Some are usable as solid or gas but majority of them are used as solutions or suspensions.

Key Points

- Solution is a homogeneous mixture of two or more substances.
- Aqueous solution is formed by dissolving substances in water.
- The component which is lesser in quantity is called solute and the component in greater quantity is called solvent.
- A solution containing less amount of solute than that is required to saturate it at a given temperature is called unsaturated solution.
- A solution that is more concentrated than that of a saturated solution is called as supersaturated solution at that particular temperature.
- Solution may be dilute or concentrated depending upon the quantity of dissolved solute in solution.
- Concentration of solutions are expressed as % w/w, % w/v, % v/w and % v/v.
- The practical unit of concentration is molarity. It is the number of moles of solute dissolved in one dm^3 of solution.
- Solubility is defined as the number of grams of the solute dissolved in 100 g of solvent to prepare a saturated solution at a given temperature . It depends upon solute-solvent interactions and temperature.
- Colloidal solutions are false solutions and in these solutions particles are bigger than in the true solutions.

EXERCISE

Multiple Choice Questions

Put a (✓) on the correct answer

1. **Mist is an example of solution:**
   (a) liquid in gas (b) gas in liquid
   (c) solid in gas (d) gas in solid

2. **Which one of the following is a `liquid in solid` solution?**
   (a) sugar in water (b) butter
   (c) opal (d) fog

3. **Concentration is ratio of:**
   (a) solvent to solute (b) solute to solution
   (c) solvent to solution (d) both a and b

4. **Which one of the following solutions contains more water?**
   (a) 2 M (b) 1 M
   (c) 0.5 M (d) 0.25 M

5. **A 5 percent (w/w) sugar solution means that:**
   (a) 5 g of sugar is dissolved in 90 g of water
   (b) 5 g of sugar is dissolved in 100 g of water
   (c) 5 g of sugar is dissolved in 105 g of water
   (d) 5 g of sugar is dissolved in 95 g of water

6. **If the solute-solute forces are strong enough than those of solute-solvent forces. The solute:**
   (a) dissolves readily (b) does not dissolve
   (c) dissolves slowly (d) dissolves and precipitates.

7. **Which one of the following will show negligible effect of temperature on its solubility?**
   (a) KCl (b) KNO_3
   (c) $NaNO_3$ (d) NaCl

8. **Which one of the following is heterogeneous mixture?**
   (a) milk (b) ink
   (c) milk of magnesia (d) sugar solution

9. **Tyndall effect is shown by:**
   (a) sugar solution (b) paints
   (c) jelly (d) chalk solution

10. Tyndall effect is due to:

(a) blockage of beam of light
(b) non-scattering of beam of light
(c) scattering of beam of light
(d) passing through beam of light

11. If 10 cm^3 of alcohol is dissolved in 100 g of water, it is called:

(a) % w/w (b) %w/v
(c) % v/w (d) %v/v

12. When a saturated solution is diluted it turns into:

(a) supersaturated solution (b) saturated solution
(c) a concentrated solution (d) unsaturated solution

13. Molarity is the number of moles of solute dissolved in:

(a) 1kg of solution (b) 100 g of solvent
(c) 1 dm^3 of solvent (d) 1 dm^3 of solution.

Short Answer Questions

1. Why suspensions and solutions do not show tyndall effect, while colloids do?
2. What is the reason for the difference between solutions, colloids and suspensions?
3. Why the suspension does not form a homogeneous mixture?
4. How will you test whether given solution is a colloidal solution or not?
5. Classify the following into true solution and colloidal solution:

   Blood, starch solution, glucose solution, toothpaste, copper sulphate solution, silver nitrate solution.
6. Why we stir paints thoroughly before using?
7. Which of the following will scatter light and why?
   sugar solution, soap solution and milk of magnesia.
8. What do you mean, like dissolves like? Explain with examples
9. How does nature of attractive forces of solute-solute and solvent-solvent affect the solubility?
10. How you can explain the solute-solvent interaction to prepare a NaCl solution?
11. Justify with an example that solubility of a salt increases with the increase in temperature.
12. What do you mean by volume/volume %?

Long Answer Questions

1. What is saturated solution and how it is prepared?
2. Differentiate between dilute and concentrated solutions with a common example.
3. Explain, how dilute solutions are prepared from concentrated solutions?
4. What is molarity and give its formula to prepare molar solution?
5. Explain the solute-solvent interaction for the preparation of solution.
6. What is general principle of solubility?
7. Discuss the effect of temperature on solubility.
8. Give the five characteristics of colloid.
9. Give at least five characteristics of suspension.

Numericals

1. A solution contains 50 g of sugar dissolved in 450 g of water. What is concentration of this solution?
2. If 60 cm^3 of alcohol is dissolved in 940 cm^3 of water, what is concentration of this solution?
3. How much salt will be required to prepare following solutions (atomic mass: K=39; Na=23; S=32; O=16 and H=1)
   a. 250 cm^3 of KOH solution of 0.5 M
   b. 600 cm^3 of $NaNO_3$ solution of 0.25 M
   c. 800 cm^3 of Na_2SO_4 solution of 1.0 M
4. When we dissolve 20 g of NaCl in 400 cm^3 of solution, what will be its molarity?
5. We desire to prepare 100 cm^3 0.4 M solution of $MgCl_2$, how much $MgCl_2$ is needed?
6. 12 M H_2SO_4 solution is available in the laboratory. We need only 500cm^3 of 0.1 ***M*** solution , how it will be prepared?

Electrochemistry

Major Concepts

7.1 Oxidation and reduction
7.2 Oxidation states and rules for assigning oxidation states
7.3 Oxidizing and reducing agents.
7.4 Oxidation – reduction reactions
7.5 Electrochemical cells
7.6 Electrochemical industries
7.7 Corrosion and its prevention

Time allocation

Teaching periods	18
Assessment periods	03
Weightage	18%

Students Learning Outcomes

Students will be able to:

- Define oxidation and reduction in terms of loss or gain of oxygen or hydrogen.
- Define oxidation and reduction in terms of loss or gain of electrons.
- Identify the oxidizing and reducing agents in a redox reaction.
- Define oxidizing and reducing agents in a redox reaction.
- Define oxidation state.
- State the common rules used for assigning oxidation numbers to free elements, ions (simple and complex), molecules, atoms .
- Determine the oxidation number of an atom of any element in a compound.
- Describe the nature of electrochemical processes.
- Sketch an electrolytic cell, label the cathode and the anode.
- Identify the direction of movement of cations and anions towards respective electrodes.
- List the possible uses of an electrolytic cell.
- Sketch a Daniel cell, labelling the cathode, the anode, and the direction of flow of the electrons.
- Describe how a battery produces electrical energy.
- Identify the half-cell in which oxidation occurs and the half-cell in which reduction occurs given a voltaic cell.
- Distinguish between electrolytic and voltaic cells.
- Describe the methods of preparation of alkali metals.
- Describe the manufacture of sodium metal from fused NaCl.
- Identify the formation of by-products in the manufacture of sodium metal from fused NaCl.

- Describe the method of recovering metal from its ore.
- Explain electrolytic refining of copper.
- Define corrosion.
- Describe rusting of iron as an example of corrosion.
- Summarize the methods used to prevent corrosion.
- Explain electroplating of metals on steel (using examples of zinc, tin and chromium plating).

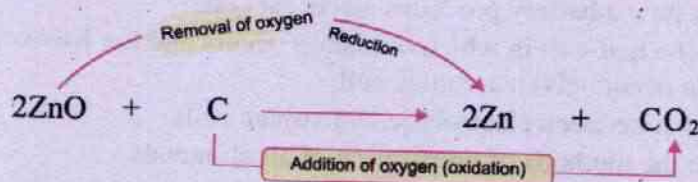
Introduction

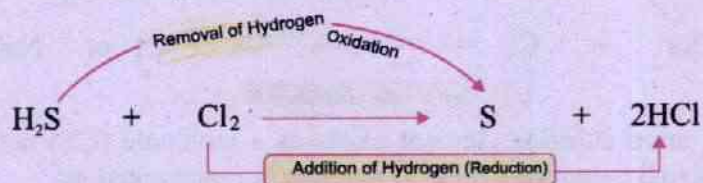
Electrochemistry is the branch of Chemistry that deals with the relationship between electricity and chemical reactions. It involves oxidation and reduction reactions, which are also known as *redox* reactions. Redox reactions either take place spontaneously and produce electricity or electricity is used to drive non-spontaneous reactions. Spontaneous reactions are those which take place on their own without any external agent. Non-spontaneous reactions are those which take place in the presence of an external agent. These reactions take place in galvanic or electrolytic cells. Electrolysis of fused sodium chloride produces sodium metal and that of brine solution produces sodium hydroxide. The corrosion process of iron along with its preventions, are discussed in detail.

7.1 OXIDATION AND REDUCTION REACTIONS

One concept of oxidation and reduction is based upon either addition or removal of oxygen or addition or removal of hydrogen in a chemical reaction. So according to this concept:

***Oxidation** is defined as addition of oxygen or removal of hydrogen during a chemical reaction.* ***Reduction** is defined as addition of hydrogen or removal of oxygen during a chemical reaction.* Both of these processes take place simultaneously in a reaction, we can say **where there is oxidation there is reduction.**

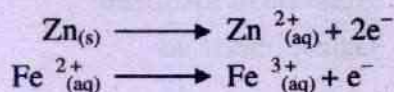
Let us first discuss an example to understand the concept based on **addition and removal of oxygen**. A reaction between zinc oxide and carbon takes place by the removal of oxygen (reduction) from zinc oxide and addition of oxygen (oxidation) to carbon. It is represented as

Let us have another example based upon **removal or addition of hydrogen.** A reaction between hydrogen sulphide and chlorine takes place by oxidation of

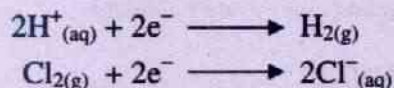
hydrogen sulphide and reduction of chlorine. Hydrogen is being removed from H_2S and added to chlorine. It is.represented as

So, *a chemical reaction in which oxidation and reduction processes are involved is called* ***oxidation- reduction reaction or redox reaction.***

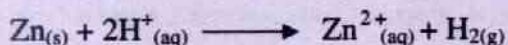
7.1.1 Oxidation and Reduction in Terms of Loss or Gain of Electron

In chemistry, there are many chemical reactions which do not involve oxygen or hydrogen, but they are considered redox reactions. To deal with these reactions, new concept 'loss or gain of electrons' is used. Therefore, reactions which involve *'loss or gain of electrons' are also called oxidation and reduction reactions.* According to this concept:

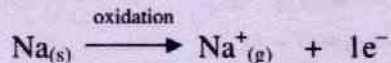
Oxidation *is loss of electrons by an atom or an ion. e.g.*

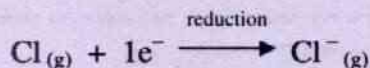
$$Zn_{(s)} \longrightarrow Zn^{2+}_{(aq)} + 2e^-$$
$$Fe^{2+}_{(aq)} \longrightarrow Fe^{3+}_{(aq)} + e^-$$

Reduction *is gain of electrons by an atom or ion. e.g.*

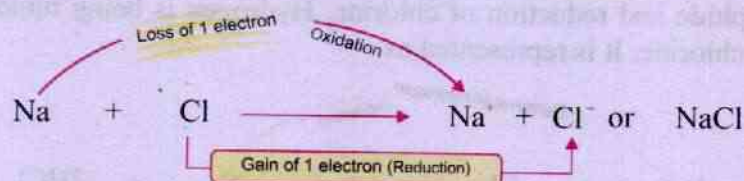
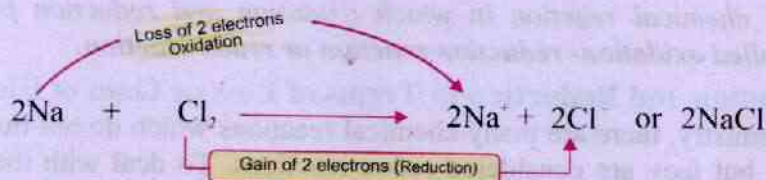
$$2H^+_{(aq)} + 2e^- \longrightarrow H_{2(g)}$$
$$Cl_{2(g)} + 2e^- \longrightarrow 2Cl^-_{(aq)}$$

The overall redox reaction is sum of both processes, written as

$$Zn_{(s)} + 2H^+_{(aq)} \longrightarrow Zn^{2+}_{(aq)} + H_{2(g)}$$

Let us have another simple example to understand this concept. A reaction between sodium metal and chlorine takes place in three steps.

First sodium atom losses an electron, to form sodium ion, such as;

$$Na_{(s)} \xrightarrow{\text{oxidation}} Na^+_{(g)} + 1e^-$$

Simultaneously, this electron is accepted by chlorine atom (reduction process), as chlorine atom needs one electron to complete its octet. As a result, chlorine atom changes to chloride ion. Such as;

$$Cl_{(g)} + 1e^- \xrightarrow{\text{reduction}} Cl^-_{(g)}$$

Ultimately, both these ions attract each other to form sodium chloride.

Complete redox reaction is sum of the oxidation and reduction reactions between sodium and chlorine atoms and it is represented as

Keep in mind chlorine element exists as a molecule (Cl_2) not as atoms (Cl). Therefore, the actual balanced chemical reaction is represented as:

We can summarize all these concepts as:

Oxidation	Reduction
Addition of oxygen	Removal of oxygen
Removal of hydrogen	Addition of hydrogen
Loss of electrons	Gain of electrons

Test yourself 7.1

i- *How can you justify that a reaction between magnesium and oxygen is a redox reaction, while the reaction shows only addition of oxygen (oxidation)*

$$2\,Mg + O_2 \longrightarrow 2\,MgO$$

ii. *A reaction between carbon and oxygen involved only addition of oxygen (oxidation), but, it is called a redox reaction. Comment on this.*

iii. *Oxidation and reduction proceed simultaneously. Explain, with an example.*

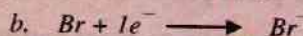
iv. *Identify which of the following is oxidation or reduction reaction*

a. $K \longrightarrow K^{+} + 1e^{-}$

b. $Br + 1e^{-} \longrightarrow Br^{-}$

c. $Cu \longrightarrow Cu^{2+} + 2e^{-}$

d. $I^{-} \longrightarrow I + 1e^{-}$

e. $Fe^{2+} \longrightarrow Fe^{3+} + 1e^{-}$

v. *An element M reacts with another element X to form MX_2. In terms of loss or gain of electrons, identify the element which is oxidized and which is reduced.*

vi. *How can you justify that the following reaction is not only an oxidation reaction but also a complete redox reaction.*

$$FeO + CO \longrightarrow Fe + CO_2$$

vii. *Explain the term oxidation on the basis of electronic concept with an example*

7.2 OXIDATION STATE AND RULES FOR ASSIGNING OXIDATION STATE

Oxidation state or **oxidation number** (*O.N.*) *is the apparent charge assigned to an atom of an element in a molecule or in an ion*. For example in HCl, oxidation number of H is + 1 and that of Cl is –1.

Rules for assigning oxidation numbers (O.N.)

i. The oxidation number of all elements in the free state is zero.
ii. The oxidation number of an ion consisting of a single element is the same as the charge on the ion.
iii. The oxidation number of different elements in the periodic table is: in Group 1 it is +1, in Group 2 it is +2 and in Group 13 it is +3.
iv. The oxidation number of hydrogen in all its compounds is +1. But in metal hydrides it is –1.
v. The oxidation number of oxygen in all its compounds is –2. But it is –1 in peroxides and +2 in OF_2
vi. In any substance the more electronegative atom has the negative oxidation number.
vii. In neutral molecules, the algebraic sum of the oxidation numbers of all the elements is zero.
viii. In ions, the algebraic sum of oxidation number equals the charge on the ion.

> **Remember!**
> *It is important to note that while **assigning oxidation numbers** the sign precedes the number. It is written as +2. Whereas, the apparent charge on an atom, ion or molecule which is called **valency,** is written as the sign followed by the number i.e 2+.*

Example 7.1

Find oxidation number of nitrogen in HNO_3 when the oxidation numbers of H = +1 and O = –2.

Solution

By applying formula in compound, sum of all oxidation numbers is zero. In case of this compound (HNO_3) it becomes:

[*O.N.* of H] + [*O.N.* of N] + 3[*O.N.* of O] = 0

Putting the values in above formula

$$[+1] + [O.N.\text{ of N}] + 3[-2] = 0$$
$$+1 + O.N.\text{ of N} + [-6] = 0$$
or $$O.N.\text{ of Nitrogen} = 6 - 1 = +5$$

Example 7.2

Calculate the oxidation number of sulphur in H_2SO_4, when O.N. of H = +1 and O.N. of O = −2.

Solution

Applying the formula for H_2SO_4,

$$2[O.N.\text{ of H}] + [O.N.\text{ of S}] + 4[O.N.\text{ of O}] = 0$$

Putting the values in above formula

$$2[+1] + [O.N.\text{of S}] + 4[-2] = 0$$
$$2 + [O.N.\text{ of S}] - 8 = 0$$
$$O.N.\text{ of S} = 8 - 2 = +6$$

Example 7.3

Find out the oxidation number of chlorine in $KClO_3$.

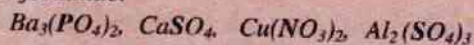
As *O.N.* of K = +1 and *O.N.* of O = −2

Solution

Putting the values in formula, we get

$$[+1] + [O.N.\text{of Cl}] + 3[-2] = 0$$
$$[+1] + [O.N.\text{ Cl}] + [-6] = 0$$
$$O.N.\text{ of Cl} = 6-1$$
$$= +5$$

Test yourself 7.2

- i- *Find out the oxidation numbers of the following elements marked in* **bold** *in the formulae:*
  $Ba_3(PO_4)_2$, $CaSO_4$, $Cu(NO_3)_2$, $Al_2(SO_4)_3$
- ii- *In a compound MX_3, find out the oxidation number of M and X.*
- iii- *Why the oxidation number of oxygen in OF_2 is +2*
- iv- *In H_2S, SO_2 and H_2SO_4 the sulphur atom has different oxidation number. Find out the oxidation number of sulphur in each compound.*
- v. *An element X has oxidation state 0. What will be its oxidation state when it gains three electrons?*
- vi. *An element in oxidation state +7 gains electrons to be reduced to oxidation state +2. How many electrons did it accept?*
- vii. *If the oxidation state of an element changes from +5 to -3. Has it been reduced or oxidized? How many electrons are involved in this process?*

7.3 OXIDIZING AND REDUCING AGENTS

An **oxidizing agent** *is the specie that oxidises a substance by taking electrons from it.* The substance (atom or ion) *which is* reduced itself by gaining electrons is also called oxidizing agent. Non-metals are oxidizing agents because they accept electrons being more electronegative elements.

Reducing agent *is the specie that reduces a substance by donating electron to it.* The substance (atom or ion) which *is oxidized* by losing electrons is also called reducing agent. Almost all metals are good reducing agents because they have the

tendency to lose electrons. An outline of oxidation and reduction processes is given below.

***Oxidation** is 'losing electrons in a chemical reaction'*
***Reduction** is 'gaining electrons in a chemical reaction'*

***Reducing agent** – is a substance that oxidizes itself and reduces other.*
***Oxidizing agent** – is a substance that reduces itself and oxidizes other.*

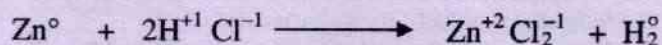
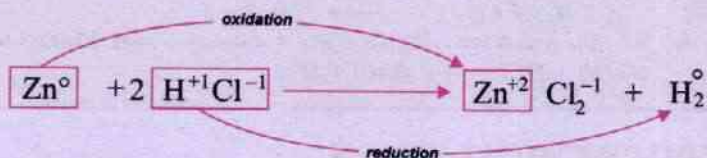
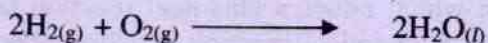
7.4 OXIDATION-REDUCTION REACTIONS

Chemical reactions in which the oxidation state of one or more substances changes are called **oxidation-reduction** or **redox** reactions. Following are the examples of redox reactions. Each reaction system consists of oxidizing and reducing agents.

Let us discuss a reaction of zinc metal with hydrochloric acid:

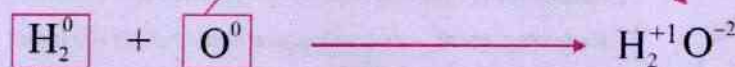
$$Zn_{(s)} + 2HCl_{(aq)} \longrightarrow ZnCl_{2(aq)} + H_{2(g)}$$

The oxidation states or oxidation numbers of all the atoms or ions in this reaction are indicated below:

$$Zn^{\circ} + 2H^{+1}Cl^{-1} \longrightarrow Zn^{+2}Cl_2^{-1} + H_2^{\circ}$$

Let us find out the atoms that are oxidized or reduced or whether there is a change in their oxidation state, it is indicated as follows:

Similarly in the case of formation of water from hydrogen and oxygen gases, redox reaction takes place as follows:

$$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(l)}$$

The oxidation states or oxidation numbers of all the atoms or ions in this reaction are:

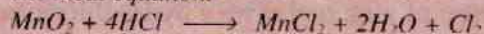
$$2H_2^{\circ} + O_2^{\circ} \longrightarrow 2H_2^{+1}O^{-2}$$

Let us find out the atoms that are oxidized or reduced in this reaction; with the help of figure below:

Test yourself 7.3

i. *In the following reaction, how can you justify that H_2S is oxidized and SO_2 is reduced.*

$$SO_2 + 2H_2S \longrightarrow 2H_2O + 3S$$

ii. *The reaction between MnO_2 and HCl is a redox reaction written as balance chemical equation.*

$$MnO_2 + 4HCl \longrightarrow MnCl_2 + 2H_2O + Cl_2$$

Find out:

a. *The substance oxidized.*
b. *The substance reduced.*
c. *The substance which acts as an oxidizing agent.*
d. *The substance which acts as a reducing agent.*

iii. *The following reactions are redox reactions.*
Find out the element which has been reduced and the element which has been oxidized.

a. $Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu$
b. $Cu + 2AgNO_3 \longrightarrow Cu(NO_3)_2 + 2Ag$
c. $H_2S + Cl_2 \longrightarrow 2HCl + S$

iv. *Why the following reaction is not a redox reaction. Explain with reasons?*

$$NaOH + HCl \longrightarrow NaCl + H_2O$$

7.5 ELECTROCHEMICAL CELLS

Electrochemical cell *is a system in which two electrodes are dipped in the solution of an electrolyte which are connected to the battery.* Electrochemical cell is an energy storage device in which either a chemical reaction takes place by using electric current (electrolysis) or chemical reaction produces electric current(electric conductance).

Electrochemical cells are of two types.

i. Electrolytic cells ii. Galvanic cells

7.5.1 Concept of Electrolytes

The substances, which can conduct electricity in their aqueous solutions or molten states, are called **electrolytes**. For example, solutions of salts, acids or bases are good electrolytes. The electricity cannot pass through solid NaCl but in aqueous solution

and in molten state, it does conduct. Electrolytes are classified into two groups depending upon their extent of ionization in solution.

7.5.1.1 Strong Electrolytes

The electrolytes which ionize almost completely in their aqueous solutions and produce more ions, are called **strong electrolytes.** Example of strong electrolytes are aqueous solutions of NaCl, NaOH and H_2SO_4 etc.

$$NaOH_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + OH^-_{(aq)}$$

7.5.1.2 Weak Electrolytes

The electrolytes which ionize to a small extent when dissolved in water and could not produce more ions are called **weak electrolytes**. Acetic acid (CH_3COOH) and $Ca(OH)_2$ when dissolved in water, ionize to a small extent and are good examples of weak electrolytes. Weak electrolytes do not ionize completely. For example, ionization of acetic acid in water produces less ions:

$$CH_3COOH_{(l)} + H_2O_{(l)} \rightleftharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

As a result the weak electrolyte is a poor conductor of electricity.

7.5.1.3 Non-Electrolytes

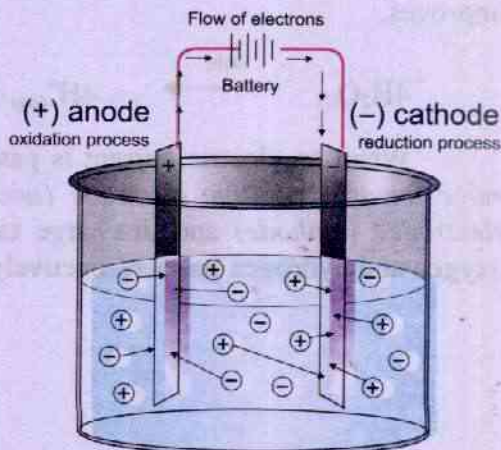
The substances, which do not ionize in their aqueous solutions and do not allow the current to pass through their solutions, are called **non-electrolytes**. For example, sugar solution and benzene are non-electrolytes.

7.5.2 Electrolytic cells

The type of electrochemical cell in which a non-spontaneous chemical reaction takes place when electric current is passed through the solution, is called an **electrolytic cell.** The process that takes place in an electrolytic cell is called electrolysis. It is *defined as the chemical decomposition of a compound into its components by passing current through the solution of the compound or in the molten state of the compound.* Examples of these cells are Downs cell, Nelson's cell.

Fig. 7.1 Electrolytic cell

7.5.2.1 Construction of an Electrolytic Cell

An electrolytic cell consists of a solution of an electrolyte, two electrodes

(anode and cathode) that are dipped in the electrolytic solution and connected to the battery. The electrode connected to positive terminal is called anode and electrode connected to the negative terminal is called cathode as shown in figure 7.1.

7.5.2.2 Working of an Electrolytic Cell

When electric current is applied from battery, the ions in the electrolyte migrate to their respective electrodes. The anions, which are negatively charged, move towards the anode and discharge there by losing their electrons. Thus *oxidation takes place at anode.* While cations, which are positively charged ions, move towards cathode. *Cations gain electrons from the electrode and as a result reduction takes place at cathode.* For example, when fused salt of sodium chloride is electrolysed the following reactions take place during this process:

$$2\,NaCl_{(s)} \longrightarrow 2\,Na^{+}_{(l)} + 2\,Cl^{-}_{(l)}$$

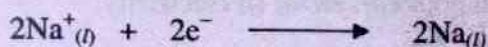
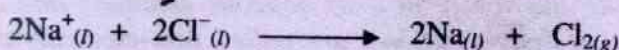
Oxidation reaction at anode:

$$2Cl^{-}_{(l)} \longrightarrow Cl_{2(g)} + 2e^{-}$$

Reduction reaction at cathode:

$$2Na^{+}_{(l)} + 2e^{-} \longrightarrow 2Na_{(l)}$$

Overall reaction:

$$2Na^{+}_{(l)} + 2Cl^{-}_{(l)} \longrightarrow 2Na_{(l)} + Cl_{2(g)}$$

7.5.2.3 Electrolysis of Water

Pure water is a very weak electrolyte. It ionizes to a very small extent. The concentrations of hydrogen ions (H^+) and hydroxyl ions (OH^-) are both at 10^{-7} *mol dm*$^{-3}$ respectively. When a few drops of an acid are added in water, its conductivity improves.

$$4H_2O_{(l)} \xrightarrow{\text{acid}} 4H^{+}_{(aq)} + 4OH^{-}_{(aq)}$$

When an electric current is passed through this acidified water, OH^- **anions** *move towards positive electrode (anode)* and H^+ **cations** *move towards negative electrode (cathode)* and discharge takes place at these electrodes. They produce oxygen and hydrogen gases respectively at anode and cathode as shown in figure 7.2.

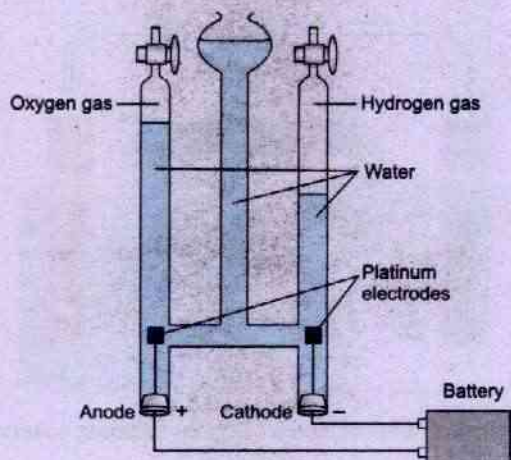



Fig. 7.2 Electrolytic cell showing electrolysis of water

The redox reaction taking place in the electrolytic bath can be shown as following:

Oxidation reaction at anode:

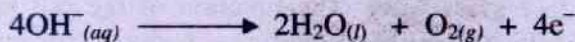
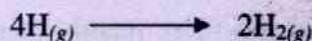
$$4OH^-_{(aq)} \longrightarrow 2H_2O_{(l)} + O_{2(g)} + 4e^-$$

Reduction reaction at cathode:

$$4H^+_{(aq)} + 4e^- \longrightarrow 4H_{(g)}$$

$$4H_{(g)} \longrightarrow 2H_{2(g)}$$

Overall reaction:

$$2H_2O_{(l)} \longrightarrow 2H_{2(g)} + O_{2(g)}$$

7.5.3. Galvanic Cell

The electrochemical cell in which a spontaneous chemical reaction takes place and generates electric current is called **galvanic or voltaic cell**. Example of this type of cell is a Daniel cell.

A. Volta (1745-1827) was an Italian physicist known especially for the development of the first electric cell in 1800

7.5.3.1 Construction of a Daniel Cell

A galvanic cell consists of two cells, each called as **half-cell,** connected electrically by a salt-bridge. In each of the half-cell, an electrode is dipped in 1*M* solution of its own salt and connected through a wire to an external circuit. Figure.7.3 shows a typical galvanic cell.

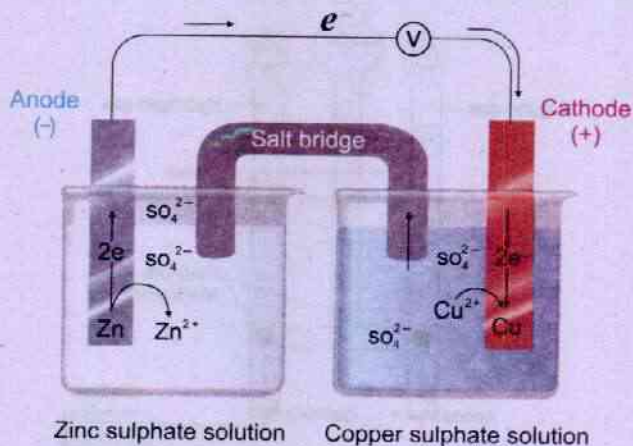



Fig. 7.3 A Daniel Cell

The left half-cell consists of an electrode of zinc metal dipped in 1 *M* solution of zinc sulphate. The right half cell is a copper electrode dipped in I*M* solution of copper sulphate. Salt bridge is a U shaped glass tube. It consists of saturated solution of strong electrolyte supported in a jelly type material. The ends of the U tube are sealed with a porous material like glass wool. The function of the salt bridge is to keep the solutions of two half cells neutral by providing a pathway for migration of ions.

7.5.3.2 Working of the Cell

The Zn metal has tendency to lose electrons more readily than copper. As a result oxidation takes place at Zn-electrode. The electrons flow from Zn-electrode through the external wire in a circuit to copper electrode. These electrons are gained by the copper ions of the solution and copper atoms deposit at the electrode. The respective oxidation and reduction processes going on at two electrodes are as follows:

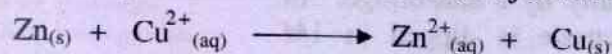
Half-cell reaction at anode (oxidation)

$$Zn_{(s)} \longrightarrow Zn^{2+}{}_{(aq)} + 2e^-$$

Half-cell reaction at cathode (reduction)

$$Cu^{2+}{}_{(aq)} + 2e^- \longrightarrow Cu_{(s)}$$

Overall galvanic reaction is the sum of these two half-cell reactions

$$Zn_{(s)} + Cu^{2+}{}_{(aq)} \longrightarrow Zn^{2+}{}_{(aq)} + Cu_{(s)}$$

As a result of redox reaction, electric current is produced. The batteries which are used for starting automobiles, running calculators and toys and to lit the bulbs work on the same principle.

A Comparison of Electrolytic and Galvanic Cells

Electrolytic cell	Galvanic cell
i. It consists of one complete cell, connected to a battery.	It consist of two half_cells connected through a salt bridge.
ii. Anode has positive charge while cathode has negative charge	Anode has negative charge while cathode has positive charge.
iii. Electrical energy is converted into chemical energy	Chemical energy is converted into electrical energy
iv. Current is used for a non-spontaneous chemical reaction to take place.	Redox reaction takes place spontaneously and produces electric current

Test yourself 7.4

i. Why are the strong electrolytes termed as good conductors?
ii. Does non-electrolytes forms ions in solution?
iii. What is difference between a strong electrolyte and a weak electrolyte?
iv. Identify a strong or weak electrolyte among the following compounds: $CuSO_4$, H_2CO_3, $Ca(OH)_2$, HCl, $AgNO_3$
v. Which force drives the non-spontaneous reaction to take place?
vi. Which type of chemical reaction takes place in electrolytic cell?
vii. What type of reaction takes place at anode in electrolytic cell?
viii. Why the positively charged electrode is called anode in electrolytic cell?
ix. In the electrolysis of water, towards which terminal H^+ ions move?
x. In the electrolysis of water, where is the oxygen produced?
xi. Towards which electrode of the electrolytic cell moves the cations and what does they do there?
xii. How the half-cells of a galvanic cell are connected? What is function of salt bridge?

7.6 ELECTROCHEMICAL INDUSTRIES

7.6.1 Manufacture of Sodium Metal from Fused NaCl

On the industrial scale molten sodium metal is obtained by the electrolysis of fused NaCl in the **Downs cell.** This electrolytic cell is a circular furnace. In the center there is a large block of graphite, which acts as an anode while cathode around it is made of iron as shown in figure 7.4.

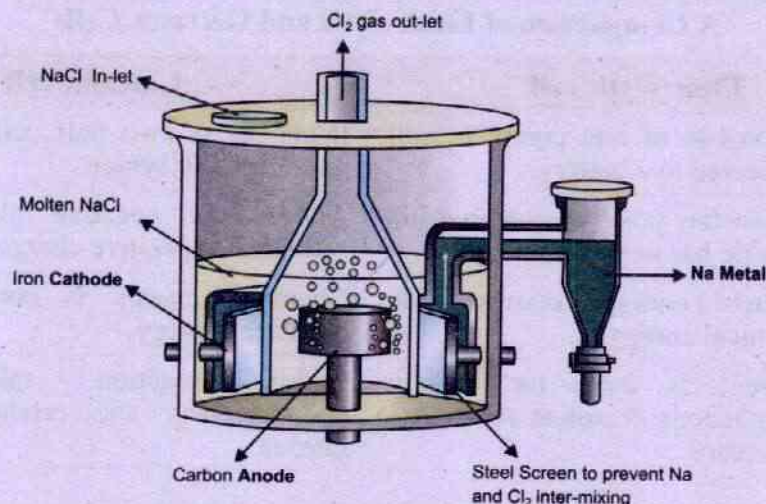
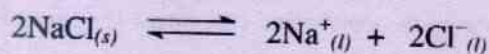



Fig. 7.4 Downs Cell for production of Sodium Metal

7.6.1.1 Working of Downs Cell

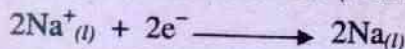
The fused NaCl produces Na^+ and Cl^- ions, which migrate to their respective electrodes on the passage of electric current. The electrodes are separated by steel gauze to prevent the contact between the products. The Cl^- ions are oxidized to give Cl_2 gas at the anode. It is collected over the anode within an inverted cone-shaped structure. While Na^+ are reduced at cathode and molten Na metal floats on the denser molten salt mixture from where it is collected in a side tube. Following reactions take place during the electrolysis of the molten sodium chloride:

Molten NaCl ionizes as:

$$2NaCl_{(s)} \rightleftharpoons 2Na^+_{(l)} + 2Cl^-_{(l)}$$

Half-cell reaction at anode (oxidation)

$$2Cl^-_{(l)} \longrightarrow Cl_{2(g)} + 2e^-$$

Half-cell reaction at cathode (reduction)

$$2Na^+_{(l)} + 2e^- \longrightarrow 2Na_{(l)}$$

Overall reaction is the sum of these two half-cell reactions

$$2NaCl_{(fused)} \longrightarrow Cl_{2(g)} + 2Na_{(l)}$$

7.6.2 Manufacture of NaOH from Brine

On industrial scale caustic soda (sodium hydroxide) NaOH, is produced in Nelson's cell by the electrolysis of aqueous solution of NaCl called brine. The schematic diagram of the cell is shown in figure 7.5. It consists of a steel tank in which graphite anode is suspended in the center of a U shaped perforated iron cathode. This iron cathode is internally lined with asbestos diaphragm. Electrolyte brine is present inside the iron cathode.

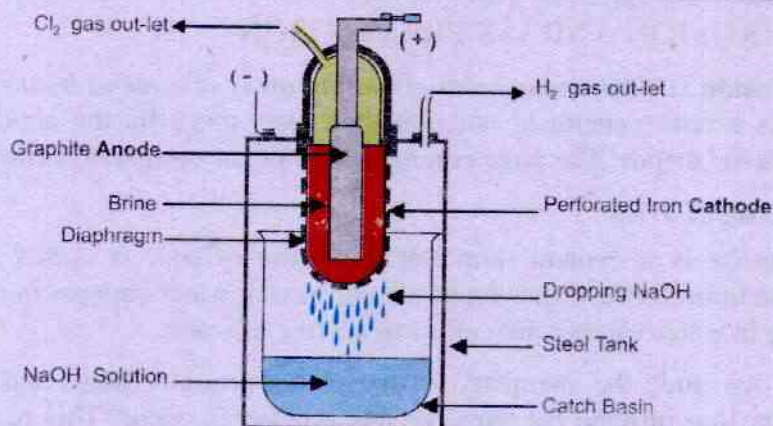



Fig. 7.5 Nelson's Cell for production of NaOH

Working of Nelson's Cell

Aqueous solution of sodium chloride consists of Na^+, Cl^-, H^+ and OH^- ions. These ions move towards their respective electrodes and redox reactions take place at these electrodes. When electrolysis takes place Cl^- ions are discharged at anode and Cl_2 gas rises into the dome at the top of the cell. The H^+ ions are discharged at cathode and H_2 gas escapes through a pipe. The sodium hydroxide solution slowly percolates into a catch basin.

Brine ionizes to produce ions:

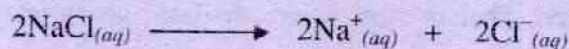
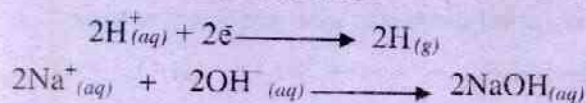
$$2NaCl_{(aq)} \longrightarrow 2Na^+_{(aq)} + 2Cl^-_{(aq)}$$

Reaction at anode (oxidation):

$$2Cl^-_{(aq)} \longrightarrow Cl_{2(g)} + 2e^-$$

Reaction at cathode (reduction):

$$2H^+_{(aq)} + 2e^- \longrightarrow 2H_{(g)}$$

$$2Na^+_{(aq)} + 2OH^-_{(aq)} \longrightarrow 2NaOH_{(aq)}$$

Overall cell reaction of this process:

$$2NaCl_{(aq)} + 2H_2O_{(l)} \longrightarrow H_{2(g)} + Cl_{2(g)} + 2NaOH_{(aq)}$$

Test yourself 7.5

i. *Anode of Downs cell is made of a non-metal, what is its name? What is the function of this anode?*
ii. *Where does the sodium metal is collected in Downs cell?*
iii. *What is the name of the by-product produced in the Downs cell?*
iv. *Are anodes of Downs cell and Nelson cell made of same element? If yes, what is its name?*
v. *What is the shape of cathode in Nelson's cell? Why is it perforated?*
vi. *Which are ions discharged at cathode in Nelson's cell and what is produced at cathode?*

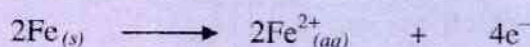
7.7 CORROSION AND ITS PREVENTION

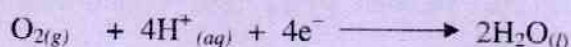
Corrosion *is slow and continuous eating away of a metal by the surrounding medium.* It is a redox chemical reaction that takes place by the action of air and moisture with the metals. The most common example of corrosion is rusting of iron.

7.7.1 Rusting of Iron

Corrosion is a general term but *corrosion of iron is called* ***rusting***. The important condition for rusting is moist air (air having water vapours in it). There will be no rusting in water vapours free of air or air free of water.

Now we study the chemistry of the rusting process. Stains and dents on the surface of the iron provide the sites for this process to occur. This region is called anodic region and following oxidation reaction takes place here:

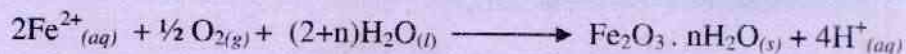
$$2Fe_{(s)} \longrightarrow 2Fe^{2+}_{(aq)} + 4e^-$$

This loss of electrons damage the object. The free electrons move through iron sheet ,until they reach to a region of relatively high O_2 concentration near the surface surrounded by water layer as shown in figure 7.6. This region acts as cathode and electrons reduce the oxygen molecule in the presence of H^+ ions:

$$O_{2(g)} + 4H^+_{(aq)} + 4e^- \longrightarrow 2H_2O_{(l)}$$

The H^+ ions are provided by the carbonic acid, which is formed because of presence of CO_2 in water. That is why acidic medium accelerates the process of rusting.

The overall redox process is completed without the formation of rust.

$$2Fe_{(s)} + O_{2(g)} + 4H^+_{(aq)} \longrightarrow 2Fe^{2+}_{(aq)} + 2H_2O_{(l)}$$

The Fe^{2+} formed spreads through out the surrounding water and react with O_2 to form the salt $Fe_2O_3.nH_2O$ which is called **rust**. It is also a redox reaction.

$$2Fe^{2+}_{(aq)} + \frac{1}{2}O_{2(g)} + (2+n)H_2O_{(l)} \longrightarrow Fe_2O_3.nH_2O_{(s)} + 4H^+_{(aq)}$$

The rust layer of iron is porous and does not prevent further corrosion. Thus rusting continues until the whole piece of iron is eaten away

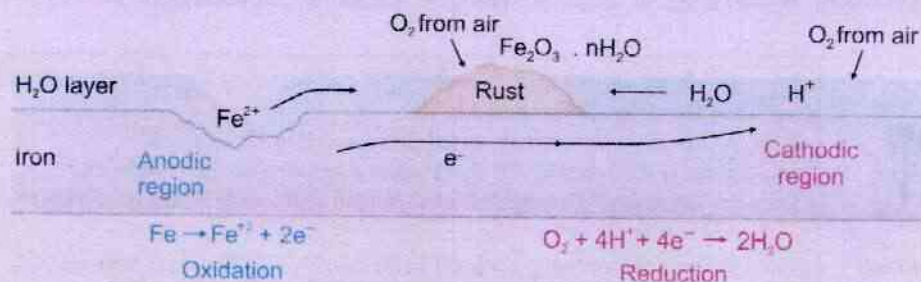



Fig. 7.6 Rusting of iron.

Does Aluminium Rust?

Aluminium corrodes but it does not rust. Rust refers only to iron and steel corrosion. A very hard material aluminium oxide protects the aluminium from further corrosion. In comparison to that when iron corrodes, its color changes and produces large red flakes known as rust. Unlike aluminium oxide, the expanding and flaking of rust exposes new metal surface to further rusting.

7.7.2 Prevention of Corrosion

7.7.2.1 Removal of stains

The regions of stains in an iron rod act as the site for corrosion. If the surface of iron is properly cleaned and stains are removed, it would prevent rusting.

7.7.2.2 Paints and greasing

Greasing or painting of the surface can prevent the rusting of iron. With development of technologies, modern paints contain a combination of chemicals called stabilizers that provide protection against the corrosion in addition to prevention against the weathering and other atmospheric effects.

7.7.2.3 Alloying

Alloy *is a homogeneous mixture of one metal with one or more other metals or non-metals.* Alloying of iron with other metals has proved to be very successful technique against rusting. The best example of alloying is the 'stainless steel', which is a good combination of iron, chromium and nickel.

7.7.2.4 Metallic coating

The best method for protection against the corrosion of metals exposed to acidic conditions is coating the metal with other metal. Corrosion resistant metals like Zn, Sn and Cr are coated on the surface of iron to protect it from corrosion. It is the most widely applied technique in the food industry where food is 'tin-packed'. The containers of iron are coated with tin to give it a longer life. Metallic coating can take place by physical as well as electrolytic methods.

1- Physical Methods (galvanizing and tin coating)

a. Zinc coating or Galvanizing

The process of coating a thin layer of zinc on iron is called galvanizing. This process is carried out by dipping a clean iron sheet in a zinc chloride bath and then heating it. After this iron sheet is removed, rolled into molten zinc metal bath and finally air-cooled. Advantage of galvanizing is that zinc protects the iron against corrosion even after the coating surface is broken.

b. Tin Coating

It involves the dipping of the clean sheet of iron in a bath of molten tin and then passing it through hot rollers. Such sheets are used in the beverage and food cans. The tin protects the iron only as long as its protective layer remains intact. Once it is broken and the iron is exposed to the air and water, a galvanic cell is established and iron rusts rapidly.

Test yourself 7.6

i. What is the difference between corrosion and rusting?
ii. What happens to iron in the rusting process?
iii. Rusting completes in how many redox reactions?
iv. Explain the role of O_2 in rusting?
v. State the best method for protection of metal from corrosion.
vi. What do you mean by galvanizing?
vii. What is the advantage of galvanizing?
viii. Why tin plated iron is rusted rapidly when tin layer is broken?
ix. Name the metal which is used for galvanizing iron?

2- Electrolytic method (Electroplating)

Electroplating *is depositing of one metal over the other by means of electrolysis.* This process is used to protect metals against corrosion and to improve their appearance. Principle of electroplating is to establish an electrolytic cell in which anode is made of the metal to be deposited and cathode of the object on which metal is to deposit. The electrolyte is in aqueous solution of a salt of the respective metal.

Procedure for Electroplating

In this process the object to be electroplated is cleaned with sand, washed with caustic soda solution and finally it is thoroughly washed with water. The anode is made of the metal, which is to be deposited like Cr, Ni. The cathode is made up of the object that is to be electroplated like some sheet made up of iron. The electrolyte in this system is a salt of the metal being deposited. The electrolytic tank is made of cement, glass or wood in which anode and cathode are suspended. These electrodes are connected with a battery. When the current is passed, the metal from anode dissolves in the solution and metallic ions migrate to the cathode and discharge or deposit on the cathode (object). As a result of this discharge, a thin layer of metal deposits on the object, which is then pulled out and cleaned. Some examples of electroplating are discussed here.

a. Electroplating of Silver

The electroplating of silver is carried out by establishing an electrolytic cell. The pure piece of silver strip acts as anode that is dipped in silver nitrate solution. The cathode is the metallic object to be coated such as spoon. When the current

is passed through the cell, the anode dissolves to produce Ag^+ ions, that migrate towards the cathode. At cathode they are discharged and deposited on the object e.g. spoon. The chemical reaction can be represented as:

At anode: $Ag_{(s)} \longrightarrow Ag^+_{(aq)} + e^-$

At cathode: $Ag^+_{(aq)} + e^- \longrightarrow Ag_{(s)}$

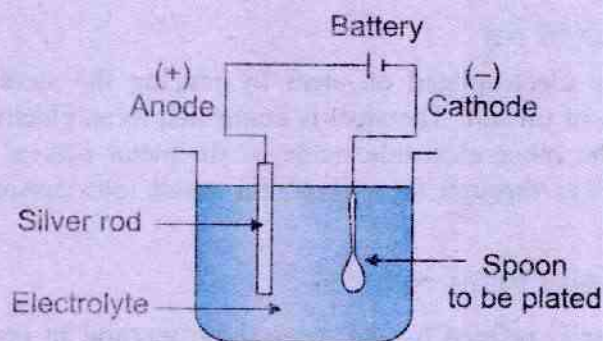
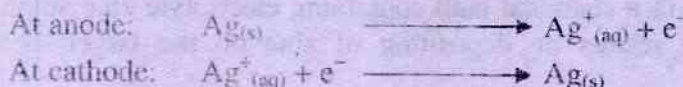



Fig. 7.7 Electroplating of an object.

Common examples of silver plating are tablewares, cutlery, jewelry and steel objects.

b. Electroplating of Chromium

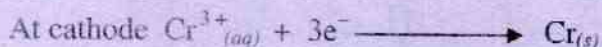
The electroplating of chromium is carried out in the same way as that of silver. The object to be electroplated is dipped in aqueous solution of chromium sulphate containing a little sulphuric acid, that acts as an electrolyte. The object to be electroplated acts as cathode while anode is made of antimonial-lead. The electrolyte ionizes and provides Cr^{3+} ions, which reduce and deposit at cathode.

Electrolyte produces the following ions:

$$Cr_2(SO_4)_{3(s)} \xrightarrow{\text{water}} 2Cr^{3+}_{(aq)} + 3SO_4^{2-}{}_{(aq)}$$

Reactions at the electrodes are as follows:

At anode $4OH^-_{(aq)} \longrightarrow 2H_2O_{(l)} + O_{2(g)} + 4e^-$

At cathode $Cr^{3+}_{(aq)} + 3e^- \longrightarrow Cr_{(s)}$

For practical convenience, the steel is usually plated first with nickel or copper and then by chromium because it does not adhere well on the steel surface. Moreover, it allows moisture to pass through it and metal is stripped off. The nickel or copper provides adhesion and then chromium deposited over the adhesive layer of copper lasts longer. This type of electroplating resists corrosion and gives a bright silvery appearance to the object.

c. Electroplating of zinc

The target metal is cleaned in alkaline detergent solutions and it is treated with acid, in order to remove any rust or surface scales. Then, the zinc is deposited on the metal by immersing it in a chemical bath containing electrolyte zinc sulphate. A current is applied, which results in depositing of zinc on the target metal i.e. cathode.

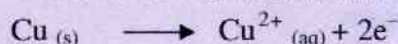
d. Electroplating of tin

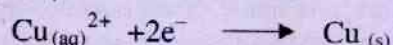
Tin is usually electroplated on steel by placing the steel into a container containing a solution of tin salt. The steel is connected to an electrical circuit, acting as cathode. While the other electrode made of tin metal acts as anode. When an electrical current passes through the circuit, tin metal ions present in the solution deposit on steel.

e. Electrolytic refining of copper

Impure copper is refined by the electrolytic method in the electrolytic cell. Impure copper acts as anode and a pure copper plate acts as cathode as shown in figure 7.8. Copper sulphate solution in water is used as an electrolyte.

Oxidation reaction takes place at the anode. Copper atoms from the impure copper lose electrons to the anode and dissolve in solution as copper ions:

$$Cu_{(s)} \longrightarrow Cu^{2+}_{(aq)} + 2e^-$$

Reduction reaction takes place at the cathode. The copper ions present in the solution are attracted to the cathode. Where they gain electrons from the cathode and become neutral and deposit on the cathode.

$$Cu_{(aq)}^{2+} + 2e^- \longrightarrow Cu_{(s)}$$

In the process impure copper is eaten away and purified copper atoms deposit on the cathode.

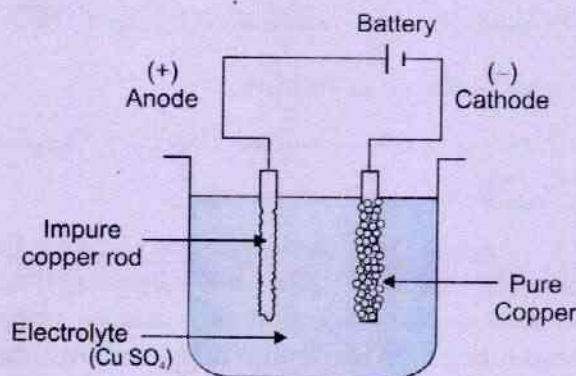



Fig. 7.8 Refining of copper in an electrolytic cell.

Test yourself 7.7

i. *Define electroplating?*
ii. *How electroplating of zinc is carried out?*
iii. *Which material is used to make cathode in electroplating?*
iv. *Why is the anode made up of a metal to be deposited during electrolysis?*

A COMPARISON OF EFFECT OF Al_2O_3 AND Fe_2O_3 FORMATION ON THEIR PARENT METALS

Aluminium has a great tendency to corrosion. However, aluminium corrosion is aluminium oxide (Al_2O_3), a very hard material that actually protects the aluminium from further corrosion. Aluminium oxide corrosion also looks a lot more like aluminium, so it isn't as easy to notice as rusted iron.

When iron corrodes the color changes and it actually expands. This expanding and color change can produce large red flakes that we all know as rust. Unlike aluminium oxide, the expanding and flaking of rust in iron exposes new metal to further rusting. That is why it is so important to provide a barrier to stop rust.

INTERACTION OF CHEMISTRY WITH PHOTOGRAPHY

In early nineteenth century photographers produced crude images using papers covered with silver nitrate or silver chloride. The exposure of light on photographic plate initiated chemical reaction. The light exposed portion became dark, depending the amount or time of exposure. That exposed plate was later on developed to show the image. Those early days "photographs" darkened with time because of ongoing chemical reaction on them. Later on procedures were developed to make the image permanent by use of mercury vapors, followed by washing with sodium hyposulfite ($Na_2S_2O_3$). It dissolved away the silver iodide from the unexposed portion of the plate, and stopped the reaction further. Although technologically more advanced, the basic procedures developed originally are still used in all silver-based photography today.

EXPLAIN HOW DECORATIVE AND PRACTICAL OBJECTS CONTAINING SILVER CAN DIFFER SIGNIFICANTLY IN THEIR PROPERTIES AND DURABILITY DEPENDING ON WHETHER THEY ARE SOLID, THICKLY PLATED WITH SILVER OR THINLY PLATED WITH SILVER.

Pure silver, also called fine silver, is relatively soft, very malleable, and easily damaged so it is commonly combined with other metals to produce a more durable product. The most popular of these alloys is sterling silver, which consists of 92.5 percent silver and 7.5 percent copper. Although any metal can make up the 7.5 percent non-silver portion of sterling, centuries of experimentation have shown copper to be its best companion, improving the metal's hardness and durability without affecting its beautiful color. The small amount of copper added to sterling has very little effect on the metal's value. Instead, the price of the silver item is affected by the labour involved in making the item, the skill of the craftsperson, and the beauty of the design. Care should also be taken to prevent silver tarnish in air, (a dulling that naturally occurs when silver reacts with sulfur or hydrogen sulfide in the ambient air.) Likewise, the art of covering a metal with other metal is also used as silver plating.

Depending upon the nature of the object, the layer of silver upon a metal is kept thick. It may be for decorative purpose of some industrial applications. Plating by silver metal has vast applications.

Key Points

- Oxidation is addition of oxygen or removal of hydrogen or loss of electrons by an element and as a result oxidation number increases.
- Reduction is addition of hydrogen or removal of oxygen or gain of electrons by an element and as a result oxidation number decreases.
- Oxidation number is the apparent charge on an atom. It may be positive or negative.
- Oxidizing agents are the species that oxidize the other element and reduce themselves. Non-metals are oxidizing agents.
- Reducing agents are species that reduce the other elements and oxidize themselves. Metals are reducing agents.
- Chemical reactions in which the oxidation state of species change are termed as redox reaction. A redox reaction involves oxidation and reduction processes taking place simultaneously.
- Redox reactions either take place spontaneously and produce energy or electricity is used to drive the reaction.
- The process in which electricity is used for the decomposition of a chemical compound is called electrolysis. It takes place in electrolytic cells such as Downs cell and Nelson's cell .
- Galvanic cells are those in which spontaneous reactions take place and generate electric current. They are also called voltaic cells.
- Sodium metal is manufactured from fused sodium chloride in the Downs cell.
- NaOH is manufactured from brine in Nelson's cell.
- Corrosion is slow and continuous eating away of a metal by the surrounding medium. The most common example of corrosion is rusting of iron.
- The rusting principle is electrochemical redox reaction, in which iron behaves as anode. Iron is oxidized to form rust $Fe_2O_3 . nH_2O$.
- Corrosion can be prevented by many methods. The most important is electroplating .
- Electroplating is depositing of one metal over the other by means of electrolysis .
- Iron can be electroplated by tin, zinc, silver or chromium.

EXERCISE

Multiple Choice Questions

Put a (✓) on the correct answer

1. **Spontaneous chemical reactions take place in:**
   (a) Electrolytic cell (b) Galvanic cell
   (c) Nelson's cell (d) Downs cell
2. **Formation of water from hydrogen and oxygen is:**
   (a) Redox reaction (b) Acid-base reaction
   (c) Neutralization (d) Decomposition
3. **Which one of the following is not an electrolytic cell?**
   (a) Downs cell (b) Galvanic cell
   (c) Nelson's cell (d) Both a and c
4. **The oxidation number of chromium in $K_2Cr_2O_7$ is:**
   (a) +2 (b) +6
   (c) +7 (d) +14
5. **Which one of the following is not an electrolyte?**
   (a) Sugar solution (b) Sulphuric acid solution
   (c) Lime solution (d) Sodium chloride solution
6. **The most common example of corrosion is:**
   (a) Chemical decay (b) Rusting of iron
   (c) Rusting of aluminium (d) Rusting of tin
7. **Nelson's cell is used to prepare caustic soda along with gases. Which of the following gas is produced at cathode:**
   (a) Cl_2 (b) H_2
   (c) O_3 (d) O_2
8. **During the formation of water from hydrogen and oxygen, which of the following does not occur:**
   (a) Hydrogen has oxidized (b) Oxygen has reduced
   (c) Oxygen gains electrons (d) Hydrogen behaves as oxidizing agent
9. **The formula of rust is:**
   (a) $Fe_2O_3.nH_2O$ (b) Fe_2O_3
   (c) $Fe(OH)_3.nH_2O$ (d) $Fe(OH)_3$
10. **In the redox reaction between Zn and HCl, the oxidizing agent is:**
    (a) Zn (b) H^+
    (c) Cl^- (d) H_2

Short Answer Questions

1. Define oxidation in terms of electrons. Give an example.
2. Define reduction in terms of loss or gain of oxygen or hydrogen. Give an example.
3. What is difference between valency and oxidation state?
4. Differentiate between oxidizing and reducing agents
5. Differentiate between strong and weak electrolytes.
6. How is electroplating of tin on steel carried out?
7. Why is steel plated with nickel before the electroplating of chromium.
8. How can you explain, that following reaction is oxidation in terms of increase of oxidation number?

$$Al^{\circ} \longrightarrow Al^{3+} + 3e^{-}$$

9. How can you prove with an example that conversion of an ion to an atom is an oxidation process?
10. Why does the anode carries negative charge in galvanic cell but positive charge in electrolytic cell? Justify with comments.
11. Where do the electrons flow from Zn electrode in Daniel's cell?
12. Why do electrodes get their names 'anode' and cathode in galvanic cell?
13. What happens at the cathode in a galvanic cell?
14. Which solution is used as an electrolyte in Nelson's cell?
15. Name the by-products produced in Nelson's cell?
16. Why is galvanizing done?
17. Why an iron grill is painted frequently?
18. Why is O_2 necessary for rusting?
19. In electroplating of chromium, which salt is used as an electrolyte?
20. Write the redox reaction taking place during the electroplating of chromium?
21. In electroplating of silver, from where Ag^+ ions come and where they deposit?
22. What is the nature of electrode used in electrolyting of chromium?

Long Answer Questions

1. Describe the rules for assigning the oxidation state
2. Find out the oxidation numbers of the underlined elements in the following compounds.

   (a) $Na_2\underline{S}O_4$ (b) $Ag\underline{N}O_3$ (c) $K\underline{Mn}O_4$

   (d) $K_2\underline{Cr}_2O_7$ (e) $H\underline{N}O_2$
3. How can a non-spontaneous reaction be carried out in an electrolytic cell? Discuss in detail.
4. Discuss the electrolysis of water.
5. Discuss the construction and working of a cell in which electricity is produced.
6. How can we prepare NaOH on commercial scale? Discuss its chemistry along with the diagram.
7. Discuss the redox reaction taking place in the rusting of iron in detail.
8. Discuss, why galvanizing is considered better than that of tin plating.
9. What is electroplating? Write down procedure of electroplating.
10. What is the principle of electroplating? How is electroplating of chromium carried out?

Chemical Reactivity

Major Concepts

1.1 Metal
1.2 Non-Metals

Time allocation	
Teaching periods	07
Assessment periods	02
Weightage	10%

Students Learning Outcomes

Students will be able to:

- Show how cation and anion are related to the terms metals and non-metals.
- Explain Alkali metals are not found in the free state in nature.
- Explain the differences in ionization energies of alkali and alkaline earth metals.
- Describe position of sodium metal in the periodic table its simple properties and uses.
- Position of calcium and magnesium in the periodic table, their simple properties and uses.
- Differentiate between soft and hard metals (iron and sodium)
- Describe the inertness of noble metals.
- Identify commercial value of silver, gold and platinum.
- Compile some important reactions of halogens.
- Name some elements that exist in nature in uncombined form.

Introduction

The different kinds of materials around us exist in variety of forms. Things like aeroplanes, trains, building frames, automobiles or even different machines and tools, are due to different properties of various metals. The non-metals exist as gases, liquids and soft or hard solids. They occupy upper right positions in the Periodic Table. Carbon, nitrogen, phosphorus, oxygen, sulphur, most of the halogens and the noble gases are non-metals. They show a variety of chemical reactivities. They form different ionic and covalent compounds, many of which are solids or gases.

8.1 METALS

Metals are the elements (except hydrogen) which are electropositive and form cations by losing electrons. Metals can be categorized.

a. Very reactive: potassium, sodium, calcium, magnesium and aluminium.
b. Moderately reactive : zinc, iron, tin and lead.
c. Least reactive or noble : copper, mercury, silver and gold.

Some common metals and non-metals in the periodic table are shown in figure 8.1.

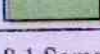
Key:

Colour of box of elements	Colour of symbol of elements
Metals Non-metals Metalloids	= Solid = Liquid = Gas

Fig. 8.1 Some common metals and non-metals.

Important **physical characteristics of metals** are listed below:

i. Almost all metals are solids (except mercury)
ii. They have high melting and boiling points. (except alkali metals)
iii. They possess metallic luster and can be polished.
iv. They are malleable (can be hammered into sheets), ductile (can be drawn into wires) and give off a tone when hit.
v. They are good conductor of heat and electricity.
vi. They have high densities.
vii. They are hard (except sodium and potassium)

Important **chemical properties** of metals are:

i. They easily lose electrons and form positive ions.
ii. They readily react with oxygen to form basic oxides.
iii. They usually form ionic compounds with non-metals.
iv. They have metallic bonding.

Do you know?

- *The most abundant metal is aluminium*
- *The most precious metal is platinum*
- *The most useable metal is iron*
- *The most reactive metal is cesium*
- *The most valuable metal is uranium*
- *The lightest metal is lithium (d = 0.53 g cm^{-3})*
- *The heaviest metal is osmium (d = 22.5 g cm^{-3})*
- *The least conductor of heat is lead.*
- *The best conductor metals are silver and gold*
- *The most ductile and malleable metals are gold and silver*

8.1.1 Electropositive Character

Metals have the tendency to lose their valance electrons. This property of a metal is termed as **electropositivity** or metallic character. The more easily a metal loses its electrons, the more electropositive it is. The number of electrons lost by an atom of a metal is called its valency. For example, sodium atom can lose 1 electron to form a positive ion

$$Na_{(s)} \longrightarrow Na^{+}_{(g)} + 1e^{-}$$

So the valency of sodium metal is 1.

Similarly zinc metal can lose 2 electrons from its valence shell. Therefore, its valency is 2.

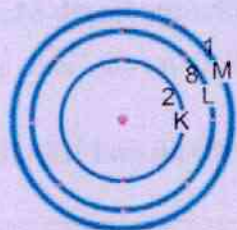
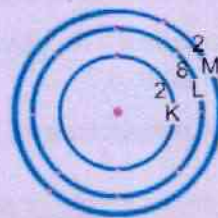
$$Zn_{(s)} \longrightarrow Zn^{2+}_{(g)} + 2e^{-}$$

Trends of electropositivity

Electropositive character increases down the group because size of atoms increases. For example, lithium metal is less electropositive than sodium which is in turn less electropositive than potassium.

Electropositive character decreases across the period from left to right in the periodic table because atomic sizes decrease due to increase of nuclear charge. It means elements at the start of a period are more metallic. This character decreases as we move from left to right along the period.

Electropositivity and ionization energy

Electropositive character depends upon the ionization energy which in turn depends upon size and nuclear charge of the atom. Small sized atoms with high nuclear charge have high ionization energy value. In this way atoms having high ionization energy are less electropositive or metallic. That is the reason alkali metals have the largest size and the lowest ionization energy in their respective periods. Therefore, they have the highest metallic character. For example, a comparison of sodium and magnesium metals is given below for understanding.

Sodium Atom
$3s^1$ electron configuration
having atomic size 186 pm,
and ionization energy 496 $kJmol^{-1}$.

Magnesium Atom
$3s^2$ electron configuration
having atomic size 160pm,
and ionization energy 738 $kJmol^{-1}$.

The 1st ionization energy of magnesium is high but the 2nd ionization energy of magnesium is very high. It becomes very difficult to remove second electron from the Mg^+ ion as nuclear charge attracts the remaining electrons strongly. As a result of this attraction size of the ion decreases.

Similarly all the elements of alkaline earth metals have high ionization energies as compared to alkali metals as shown in table 8.1.

Table 8.1 Atomic Number, Electronic Configurations and Ionization Energies (kJ/mol) of Alkali and Alkaline Earth Metals

Metal	Atomic Number	Electronic Configuration	IE	Metal	Atomic Number	Electronic Configuration	IE_1	IE_2
Li	3	[He] $2s^1$	520	Be	4	[He] $2s^2$	899	1787
Na	11	[Ne] $3s^1$	496	Mg	12	[Ne] $3s^2$	738	1450
K	19	[Ar] $4s^1$	419	Ca	20	[Ar] $4s^2$	590	1145
Rb	37	[Kr] $5s^1$	403	Sr	38	[Kr] $5s^2$	549	1064
Cs	55	[Xe] $6s^1$	377	Ba	56	[Xe] $6s^2$	503	965

Low ionization energies of alkali metals make them more reactive than alkaline earth metals.

Test yourself 8.1

i. What type of elements are metals?
ii. Name a metal which exists in liquid form?
iii. What is the nature of metal oxide?
iv. Which group of metals is highly reactive?
v. Why is sodium metal more reactive than magnesium metal?
vi. Name a metal which can be cut with knife?
vii. Name the best ductile and malleable metal?
viii. Name the metal which is the poorest conductor of heat?
ix. What do you mean by malleable and ductile?
x. Why are alkali metals more reactive than alkaline earth metals?
xi. What do you mean by metallic character?
xii. Why metallic character decreases along a period and increases in a group?

8.1.2 Comparison of Reactivities of Alkali and Alkaline Earth Metals

A comparison of physicals properties of alkali metals and alkaline earth metals is given table 8.2

Table 8.2 Comparison of Physical Properties of Alkali and Alkaline Earth Metals

Property	Sodium	Magnesium	Calcium
Appearance	Silvery white having a metallic luster, very soft and can be cut with knife	Silvery white and hard	Silvery grey and fairly harder
Atomic size, ionic size (pm)	186, 102	160, 72	197, 99
Relative density	0.98 g cm^{-3} Floats on water	1.74 g cm^{-3}	1.55 g cm^{-3}
Malleability	very malleable and ductile	Malleable and ductile	Malleable and ductile
Conductivity	Good conductor of heat and electricity	Good conductor of heat and electricity	Good conductor of heat and electricity
M.P	97 °C	650 °C	839 °C
B.P	883 °C	1090 °C	1484 °C
Ionization energy	496 kJ/mol	738, 1450 kJ/mol	590,1145 kJ/mol
Flame in air	Golden yellow	Brilliant white	Brick red

The elements in first two groups of the periodic table Group 1 and Group 2 are called '*Alkali*' and '*Alkaline earth*' metals, respectively. Alkali metals are extremely reactive elements because of their ns^1 valence shell electronic configuration. As there is only one electron in their valence shell, it can be easily given out. It is the reason that they are always found in nature as cations with +1 oxidation state. Therefore, they readily form salts with non-metals.

The alkaline earth metals atoms are smaller and have more nuclear charge. They have two electrons in their valence shells i.e. ns^2. They are also reactive but less than alkali metals.

A comparison of chemical properties and reactivities of alkali metals and alkaline earth metals is given in table 8.3

Table 8.3 Comparison of Chemical Properties and Reactivities

	Alkali Metals	Alkaline Earth Metals
1	**Occurrence**	
	They are very reactive and always occur in combined form	They are fairly reactive and also occur in combined form
2	**Electropositivity**	
	These are highly electropositive. They have ionization energy values ranging from 520 $kJmol^{-1}$ for Li to 376 $kJmol^{-1}$ for Cs.	They are less electropositive. They have ionization energy values ranging from 1757 $kJmol^{-1}$ for Be to 965 $kJmol^{-1}$ for Ba.
3	**Reaction with water**	
	They react with water vigorously at room temperature to give strong alkaline solution and hydrogen gas $2Na+2H_2O \longrightarrow 2NaOH+H_2$	They react with water less vigorously and on heating they produce weak bases $Mg + H_2O \longrightarrow MgO + H_2$ $MgO + H_2O \longrightarrow Mg(OH)_2$
4	**Reaction with O_2**	
	They immediately tarnish in air giving their oxides which form strong alkalies in water $4Na+O_2 \longrightarrow 2Na_2O$ $Na_2O+H_2O \longrightarrow 2NaOH$	They are less reactive towards oxygen and oxides are formed on heating $2Mg+O_2 \longrightarrow 2MgO$
5	**Reaction with Hydrogen**	
	They form ionic hydrides with H_2 at high temperature $2Na + H_2 \longrightarrow 2NaH$	They give hydrides under strong conditions of temperature and pressure $Ca+H_2 \longrightarrow CaH_2$
6	**Reaction with Halogens**	
	They react violently with halogens at room temperature to give halides $2Na+Cl_2 \longrightarrow 2NaCl$	They react slowly with halogen to give their halides $Ca+Cl_2 \longrightarrow CaCl_2$
7	**Reaction with Nitrogen**	
	They do not form nitrides directly	They form stable nitrides when heated with nitrogen $3Mg+N_2 \longrightarrow Mg_3N_2$
8	**Reaction with Carbon**	
	They do not react with carbon directly	They give stable carbide on heating with carbon. $Ca+2C \longrightarrow CaC_2$

Uses of sodium

i. Sodium-potassium alloy is used as a coolant in nuclear reactors.
ii. It is used to produce yellow light in sodium vapour lamps.
iii. It is used as a reducing agent in the extraction of metals like Ti.

Uses of magnesium

i. Magnesium is used in flash light bulbs and in fireworks.
ii. It is used in the manufacture of light alloys.
iii. Magnesium ribbon is used in Thermite process to ignite aluminium powder
iv. Magnesium is used as anode for prevention of corrosion.

Uses of calcium

i. It is used to remove sulphur from petroleum products.
ii. It is used as reducing agent to produce Cr, U and Zr.

Inertness of Noble Metals

The elements in which *d*-orbital are in the process of filling, constitute a group of metals called transition metals or *d*-group elements. They exhibit a variety of oxidation states. Figure 8.2 shows 'transition metals' of 4^{th}, 5^{th} and the 6^{th} period of the periodic table. There are three series of transition elements; each series consisting of ten elements.

Fig. 8.2 The Transition Elements in the Periodic Table.

Chemical behaviour of the first transition series is similar to active metals except copper. Three transition metals belonging to group 11 are copper, silver and gold. Out of them gold and silver are relatively inactive metals because they do not lose electrons easily.

Silver is white lustrous metal. It is an excellent conductor of heat and electricity. It is also highly ductile and malleable metal. Its polished surfaces are good reflectors of light. Formation of thin layer of oxide or sulphide on its surface makes it relatively unreactive. Under normal conditions of atmosphere, air does not affect

silver. It tarnishes in presence of sulphur containing compounds like H_2S.

Being very soft metal, it is rarely used as such. Alloys of silver with copper are widely used in making coins, silver-ware and ornaments. Compounds of silver are widely used in photographic films and dental preparations. Silver also has important applications in mirror industry.

Gold is a yellow soft metal. It is most malleable and ductile of all the metals. One gram of gold can be drawn into a wire of one and a half kilometre long. Gold is very non-reactive or inert metal. It is not affected by atmosphere. It is not even affected by any single mineral acid or base.

Because of its inertness in atmosphere, it is an ornamental metal as well as used in making coins. Gold is too soft to be used as such. It is always alloyed with copper, silver or some other metal.

Do you know?

Purity of gold is shown by ***carats*** *that indicates the number of parts by weight of gold that is present in 24 parts of alloy. Twenty four carat gold is pure. 22 carats gold means that 22 parts pure gold is alloyed with 2 parts of either silver or copper for making ornaments and jewelry. White gold is its alloy with palladium, nickel or zinc.*

Platinum is used to make jewelry items because of its unique characteristics like colour, beauty, strength, flexibility and resistance to tarnish. It provides a secure setting for diamonds and other gemstones, enhancing their brilliance.

An alloy of platinum, palladium and rhodium is used as catalyst in automobiles as catalytic convertor. It converts most of the toxic gases (CO, NO_2) being emitted by vehicles into less harmful carbon dioxide, nitrogen and water vapour.

Platinum is used in the production of hard disk drive coatings and fibre optic cables. Platinum is used in the manufacturing of fibre glass reinforced plastic and glass for liquid crystal displays (LCD).

Test yourself 8.2

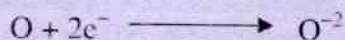
i. Give the applications of silver?
ii. Why is silver not used in pure form?
iii. What do you mean by 24 carat gold?
iv. Why is gold used to make jewelry?
v. Why is platinum used for making jewelry?
vi. What is difference between steel and stainless steel?
vii. How is platinum used as a catalyst in automobiles and what are advantages of this use?

8.2 NON-METALS

Non-metals *form negative ions (anions) by gaining electrons.* In this way, non-metals are electronegative in nature and form acidic oxides. The valency of some non-metals depend upon the number of electrons accepted by them. For example, valency of chlorine atom is 1, as it accepts only 1 electron in its outermost shell.

$$Cl + 1e^- \longrightarrow Cl^-$$

Similarly, oxygen atom can accept 2 electrons, therefore, its valency is 2.

$$O + 2e^- \longrightarrow O^{-2}$$

The non-metallic character depends upon the electron affinity and electronegativity of the atom. Small sized elements having high nuclear charge are electronegative in nature. They have high electron affinity. Therefore, they possess non-metallic character. Hence, non-metallic character decreases in a group downward and increases in a period from left to right up to halogens. Fluorine is the most non-metallic element.

The non-metals are, therefore, elements in Group-14(Carbon), Group-15 (nitrogen and phosphorus), Group-16 (oxygen, sulphur and selenium) and in Group-17 halogens (fluorine, chorine, bromine and iodine) of the periodic table. Figure 8.3 shows position of non-metals in the periodic table.

Important physical properties of non-metals are as follows:

Physical properties of non-metals change gradually but uniquely in a group of non-metals. Non-metals usually exist in all three physical states of matter. The non-metals at the top of the group are usually gases while others are either liquids or solids.

i. Solids non-metals are brittle (break easily).
ii. Non-metals are bad conductors of heat and electricity (except graphite).
iii. They are not shiny, they are dull except iodine (it is lustrous like metals).
iv. They are generally soft (except diamond).
v. They have low melting and boiling points (except silicon, graphite and diamond).
vi. They have low densities.

Fig. 8.3 The Non-Metals in Periodic Table

Important chemical properties of non-metals are as follows:

i. Their valence shells are deficient of electrons, therefore, they readily accept electrons to complete their valence shells and become stable.
ii. They form ionic compounds with metals and covalent compounds by reacting with other non-metals e.g. CO_2, NO_2, etc.
iii. Non-metals usually do not react with water.
iv. They do not react with dilute acids because non-metals are itself electron acceptors.

Electonegativity of first member of group 14, 15, 16 and 17 are higher than that of other members of the group decreasing trend of electronegativity is shown below

$$F > O > Cl > N > Br > S > C > I > P$$

8.2.1 Comparison of Reactivity of the Halogens

Elements of Group-17 of the periodic table consist of fluorine, chlorine, bromine, iodine and astatine. They are collectively called **halogens.** Fluorine and chlorine exist as diatomic gases at room temperature. Interestingly, the intermolecular forces of attraction increase downward in the group due to the increase in the size of atom. Due to this reason bromine exists as a liquid and iodine as solid. Some physical properties of halogens are shown in Table 8.4

Table 8.4 Some Physical Characteristics of Halogens

Element	Atomic No.	Electronic Configuration	Colour	Melting Point (K)	Boiling Point (K)	Electro negativities
F	9	$[He]2s^2\ 2p^5$	Pale Yellow	53	85	4.0
Cl	17	$[Ne]3s^2\ 3p^5$	Greenish yellow	172	238	3.2
Br	35	$[Ar]4s^2\ 4p^5$	Redish Brown	266	332	3.0
I	53	$[Kr]5s^2 5p^5$	Purple Black	387	457	2.7

In general their valence shell electronic configuration is $ns^2\ np^5$. Since halogens have only one electron deficit in their valence shell; either they can readily accept an electron from a metal or they can share an electron with other non-metals. Thus halogens form ionic bonds with metals and covalent bond with non-metals.

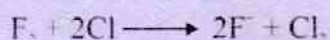
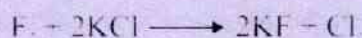
8.2.2 Important Reactions of Halogens

1) Oxidizing properties

All halogens are oxidizing agents. Fluorine is the strongest oxidizing element while iodine is the least i.e is mild oxidizing agent. Fluorine will oxidize any of halide ion (X^-) in solution and changes itself to F^- ion. Similarly chlorine will displace Br^- and I^- ions from their salt solutions and oxidize them to bromine and iodine.

$$F_2 + 2KCl \longrightarrow 2KF + Cl_2$$

$$F_2 + 2Cl \longrightarrow 2F^- + Cl_2$$

$$Cl_2 + 2KBr \longrightarrow 2KCl + Br_2$$

Solution turns from colourless to reddish brown.

Similarly, $Br_2 + 2KI \longrightarrow 2KBr + I_2$

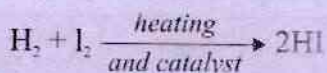
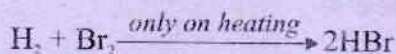
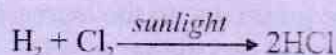
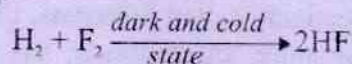
2) Reaction with hydrogen

All halogens (X_2) combine with hydrogen to give hydrogen halides (HX). However the chemical affinity for H_2 decreases down the group from F_2 to I_2. Fluorine combines with hydrogen even in the dark and cold state. Chlorine reacts with hydrogen only in the presence of sunlight. Bromine and iodine react with hydrogen only on heating.

$$H_2 + F_2 \xrightarrow[\text{state}]{\text{dark and cold}} 2HF$$

$$H_2 + Cl_2 \xrightarrow{\text{sunlight}} 2HCl$$

$$H_2 + Br_2 \xrightarrow{\text{only on heating}} 2HBr$$

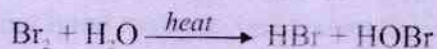
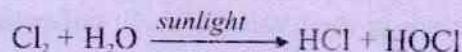
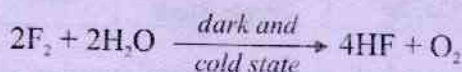
$$H_2 + I_2 \xrightarrow[\text{and catalyst}]{\text{heating}} 2HI$$

3) Reaction with water

Fluorine (F_2) decomposes water in cold state and in dark. Chlorine decomposes water in presence of sunlight. Bromine only react with water under special conditions. Iodine does not give this reaction.

$$2F_2 + 2H_2O \xrightarrow[\text{cold state}]{\text{dark and}} 4HF + O_2$$

$$Cl_2 + H_2O \xrightarrow{\text{sunlight}} HCl + HOCl$$

$$Br_2 + H_2O \xrightarrow{\text{heat}} HBr + HOBr$$

$$I_2 + H_2O \longrightarrow \text{No reaction}$$

4) Reaction with methane

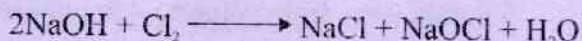
Fluorine (F_2) reacts violently with methane (CH_4) in dark while chlorine (Cl_2) does not react with methane in dark. However the presence of bright sunlight the reaction is violent.

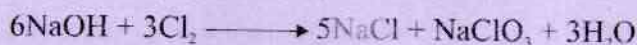
$$CH_4 + 2Cl_2 \longrightarrow C + 4HCl$$

In presence of diffused sunlight the reaction of chlorine with methane is slow and gives series of compounds i.e CH_3Cl, CH_2Cl_2, $CHCl_3$ and CCl_4.

5) Reaction with Sodium hydroxide

Chlorine reacts with cold ditute NaOH to give sodium hypochlorite and sodium chloride

$$2NaOH + Cl_2 \longrightarrow NaCl + NaOCl + H_2O$$

Cl_2 reacts with hot conc NaOH to give sodium chloride and sodium chlorate

$$6NaOH + 3Cl_2 \longrightarrow 5NaCl + NaClO_3 + 3H_2O$$

8.2.3 Significance of Non-metals

Although non-metals are fewer than metals, yet they are highly significant. They are equally important for human beings, animals and plants. In fact, life would not have been possible without the presence of non-metals on earth.

i. Major components of earth's crust, oceans and atmosphere (as shown in table 1.1) are non-metals: oxygen has the highest percentage in earth's crust (47%) and oceans (86%) and it is second (21%) to nitrogen in atmosphere. It indicates the importance of oxygen in nature. To maintain the balance for the amount of non-metals in nature, different cycles like water cycle, nitrogen cycle etc have been established naturally.

ii. Non-metals are essential part of the body structure of all living things. Human body is made up of about 28 elements. But about 96% of the mass of the human body is made up of just 4 elements i.e. oxygen 65%, carbon 18%, hydrogen 10% and nitrogen 3%. Similarly, plant bodies are made up of cellulose, which is composed of carbon, hydrogen and oxygen.

iii. Life owes to non-metals as without O_2 and CO_2 (essential gases for respiration of animals and plants respectively), life would not have been possible. In fact, these gases are essential for the **existence of life**.

iv. All eatables like carbohydrates, proteins, fats, vitamins, water, milk etc which are necessary for the growth and development of body that are made up of non-metals; carbon, hydrogen and oxygen. Its shows non-metals play a vital role for the **maintenance of life.**

v. The essential compound for the **survival of life** of both animals and plants is water, which is made up of non-metals. Water is not only the major part by mass of animals and plants bodies, but it is also essential to maintain the life. We can survive without water for days but not for a long period; its shortage may cause death.

vi. Another important non-metal is nitrogen, which is 78% in atmosphere, is necessary for the **safety of life** on earth. It controls the fire and combustion processes, otherwise all the things around us could burn with a single flame.

vii. Non-metals are playing essential role for the **communication in life.** All fossil fuels which are major source of energy; coal, petroleum and gas are made up of carbon and hydrogen. Even the essential component of combustion of fossil fuels, oxygen is also a non-metal.

viii. Non-metals protect us in a way, the clothes we wear are made of cellulose (natural fibre) or polymer (synthetic fibre).

ix. In addition to all of these, other items used in daily life such as wooden or plastic furniture, plastic sheets and bags, plastic pipes and utensils are made of non-metallic elements. Even all the pesticides, insecticides, fungicides and germicides consist of non-metals as major constituents.

Test yourself 8.3

- i. *Why valency of chlorine is 1?*
- ii. *Which factor controls the non-metallic character of the elements?*
- iii. *Why fluorine is more non-metallic than chlorine?*
- iv. *Iodine exists in solid state, can it be beaten with hammer to form sheets?*
- v. *Can liquids and gases be brittle?*
- vi. *Why the oxygen is called non-metal?*
- vii. *Name two non-metals which are both brittle and non-ductile*
- viii. *Name the most abundant non-metal in the earth's crust*
- ix. *Give the non-metallic trend in halogens*
- x. *Why do the non-metals accept electrons readily?*
- xi. *Why non-metals do not react with dilute acids while metals do react?*
- xii. *How can we distinguish a metal from a non-metal by simple physical methods?*
- xiii. *How we can distinguish a substance is metal or non-metals with the help of an acid?*
- xiv. *Why is HF a weak acid?*

Key Points

- Formation of cations of alkali and alkaline earth metals is due to their electropositive behavior.
- The chemical reactivity of alkali and alkaline earth metals, is quite different.
- Calcium and magnesium are less reactive than sodium.
- Halogens form very stable compounds with alkali metals.
- Mercury and gold exist in free elemental form in nature.

EXERCISE

Multiple Choice Questions

Put a (✓) on the correct answer.

1. **Metals can form ions carrying charges:**
   (a) Uni-positive (b) Di-positive (c) Tri-positive (d) All of them
2. **Which one of the following metal burns with a brick red flame?**
   (a) Sodium (b) Magnesium (c) Iron (d) Calcium
3. **Sodium is extremely reactive metal, but it does not react with:**
   (a) Hydrogen (b) Nitrogen (c) Sulphur (d) Phosphorus
4. **Which one of the following is the lightest metal?**
   (a) Calcium (b) Magnesium (c) Lithium (d) Sodium
5. **Pure alkali metals can be cut simply by knife but iron cannot because of alkali metals have:**
   (a) Strong metallic bonding (b) Weak metallic bonding
   (c) Non-metallic bonding (d) Moderate metallic bonding
6. **Which of the following is less malleable?**
   (a) Sodium (b) Iron (c) Gold (d) Silver
7. **Metals lose their electrons easily because:**
   (a) They are electronegative (b) They have electron affinity
   (c) They are electropositive (d) Good conductors of heat

8. **Which one of the following is brittle?**
   (a) Sodium (b) Aluminium (c) Selenium (d) Magnesium
9. **Which one of the following non-metal is lustrous?**
   (a) Sulphur (b) Phosphorus (c) Iodine (d) Carbon
10. **Non-metals are generally soft, but which one of the following is extremely hard?**
    (a) Graphite (b) Phosphorus (c) Iodine (d) Diamond
11. **Which one of the following will not react with dilute HCl?**
    (a) Sodium (b) Potassium (c) Calcium (d) Carbon

Short Answer Questions

1- Why reactivity of metals increases down the group?
2- State the physical properties of metals.
3- Why nitrogen forms compounds with alkaline earth metals directly?
4- Why the second ionization energy of magnesium is higher than the first one?
5- How oxygen reacts with group II A metals?
6- What is relationship between electropositivity and ionization energy?
7- Why electropositivity decreases from left to right in a period?
8- How electropositivity depends upon size and nuclear charge of an atom?
9- Why ionization energies of alkaline earth metals are higher than alkali metals?
10- Why are silver and gold least reactive?
11- Can pure gold be used for making ornaments? If not why?
12- Why is copper used for making electrical wires?
13- What is the trend of variation in densities of alkali metals?
14- Which metal is used for metal work?
15- Why is magnesium harder than sodium?
16- Why is calcium more electropositive than magnesium?
17- Why is ionization energy of Na less than Mg?
18- Why is the ionization energy of Na more than K?

Long Answer Questions

1- Compare and contrast the properties of alkali and alkaline earth metals.
2- Discuss the inert character of silver and gold.
3- Why are cations smaller and anions are bigger in size than their respective neutral atoms.
4- Discuss why hardness and softness of a metal depends upon its metallic bonding.
5- Give the reaction of sodium with: H_2O, O_2, Cl_2 and H_2
6- What are physical properties of calcium metal? Give its uses.
7- Write down the chemical properties of the non-metals?
8- Compare the physical properties of metals and non-metals.
9- How can you compare the softness and hardness of metals?
10- Give the chemical properties of magnesium and its uses
11- Write a comprehensive note on the electropositive character of metals?
12- Compare the ionization energies of alkali and alkaline earth metals.

ANSWERS

Chapter 1

Numericals

(1) 490 g (2) 2.41×10^{23} Ca^{2+} and 2.41×10^{23} CO_3^{2-}
(3) 9.03×10^{23} ions
(4) a. 1.55×10^{23} molecules b. 1.91×10^{23} molecules c. 1.00×10^{23} molecules
(5) a. 1.80×10^{23} ions b. 2.60×10^{23} ions c. 1.065×10^{23} ions
(6) 3.34×10^{-6} g (7) 2.87×10^{24} atoms (8) 6.17×10^{23} ions
(9) 1.65×10^{23} molecules (10) 12 g

Chapter 5

Numericals

(1) a. 1.12 atm b. 2.02 atm c. 56 cm Hg d. 126656 Pa
(2) a. 1023 K b. 423 K c. 173 °C d. 101 °C
(3) 1350 cm^3 (4) 506 mm of Hg (5) 126 °C (6) about 1:0.93
(7) 0.53 dm^3, contract (8) 30 cm^3 (9) 37.05 dm^3
(10) 1.58 atm, yes

Chapter 6

Numericals

(1) 10% m/m (2) 6% v/v (3) a- 7.0 g b- 12.75 g c- 113.6 g
(4) 0.85 M (5) 3.8 g (6) 4.16 cm^3

GLOSSARY

Absolute zero It is the temperature at which an ideal gas would have zero volume i.e. gas ceases to exist. It is represented by **K** and is equal to −273.15 °C.

Anion An anion is an atom or group of atoms having a negative charge on it.

Aqueous solution A solution which is formed by dissolving a substance in water.

Atomic mass unit (amu) is one-twelfths (1/12) the mass of an atom of carbon-12. 1 **amu** = 1.66×10^{-24} g.

Atomic number of an element is equal to the number of protons in the nucleus of each of its atom. It is represented by **Z**

Bond pairs are the electrons that pair up to form a bond.

Cations An atom or group of atoms having positive charge on it.

Chemical bond is a force of attraction between atoms that holds them together in a molecule or compound.

Chemistry is the study of composition, properties of matter, the changes that matter undergoes and the energy associated with these changes

Colloidal solutions are those in which the solute particles are larger than the true solutions but not large enough to be seen by naked eye.

Compound is a substance that which contains two or more elements (atoms) chemically combined together in a fixed ratio by mass.

Concentrated solution is that contains relatively large amount of dissolved solute in the solution.

Covalent bond is a type of bond which is formed due to mutual sharing of electrons of the atoms.

Dative or coordinate covalent bonding *is the type of covalent bond in which the bond pair of electrons is contributed by one atom only.*

Dilute solution *is that which contains relatively small amount of dissolved solute in the solution.*

Effusion *is the escape of gas molecules through a tiny hole into an evacuated space.*

Electrochemical cell *is a system in which two electrodes are dipped in the solution of an electrolyte and electrodes are connected to the battery. In this cell electric current is used to drive a non-spontaneous chemical reaction.*

Electrolysis *is the chemical decomposition of a compound into its components by passing current through the solution of the compound or the molten state of the compound.*

Electrolytes *are the substances which can conduct electricity in their aqueous solutions or molten states.*

Electron affinity *is the amount of energy released when one electron is added up in the outermost shell of an isolated gaseous atom.*

Electronegativity *is the ability of an atom to attract the shared pair of electrons towards itself in a compound.*

Electroplating *is depositing of one metal over the other by means of electric current.*

Elements *are substances which consist of only one type of atoms. They cannot be decomposed into simpler substances by any ordinary chemical process*

Empirical formula *is the simplest type of chemical formula. It shows the relative number of atoms of each element present in the compound.*

Formula unit *of ionic compounds is defined as the smallest whole number ratio of ions from which formula of a compound can be calculated.*

Free radicals *are atoms or group of atoms possessing an odd (unpaired) electron.*

Freezing point *is the temperature at which vapour pressure of a liquid becomes equal to the vapour pressure of the solid at this temperature liquid and solid coexist in dynamic equilibrium.*

Galvanic cell *It is the type of a cell in which a spontaneous chemical reaction takes place and generates electric current.*

Gram atomic mass *Atomic mass of an element expressed in grams.*

Heteroatomic molecule *consists of different kinds of atoms.*

Heterogeneous mixtures *do not have uniform composition throughout.*

Homoatomic molecule *contains same type of atoms in it.*

Homogeneous mixtures *have a uniform composition throughout.*

Ion *is an atom or group of atoms having a charge on it.*

Ionic bond *the type of chemical bond which is formed due to transference of electron from one atom to another atom.*

Ionization energy *is the amount of energy required to remove an electron from the outermost shell of an isolated gaseous atom.*

Isotopes *are the atoms of an element that have same atomic number but different `atomic mass.*

Mass number *is the sum of number of protons and neutrons in an atom. Each proton and neutron has one*

unit of mass. Mass number is represented by **'A'**.

Melting point *is defined as temperature at which solid melts and coexists in dynamic equilibrium with liquid*

Metallic bond *is a type of bond that is formed between atoms (positively charged ions) due to mobile or free electrons between them.*

Metalloids *are the elements having properties in between metals and non-metals.*

Metals *Elements which are electropositive in nature.*

Mixture *is a physical combination of two or more elements or compounds without any fixed ratio.*

Molarity *is defined as number of moles of solute dissolved in one* dm^3 *of solution. It is represented by* **'M'**

Mole *is quantity of substance that consists of fixed number of entities (atoms, molecules or formula units) equals to* 6.022×10^{23}. *This number is called* **Avogadro's number.**

Molecular compounds *The compounds that exist independently in the form of molecules.*

Molecular formula. *It shows actual number of atoms of each element in a molecule of the compound.*

Molecular ion *A molecule that has lost or gained an electron and carries a charge on it.*

Molecular mass *Sum of atomic masses of all the atoms in one molecule of a substance.*

Molecule *It is the smallest unit of an element or a compound which can exist independently.*

Monoatomic *molecule consists of only one atom.*

Non-metals *are elements which have electronegative character.*

Octet rule *The tendency of atoms to attain 8 electron configuration in the outermost shell either by gaining or by losing electrons.*

Oxidation *is defined as the loss of electron by an element or anion.*

Oxidation state *or* **oxidation number** *is the apparent charge assigned to an atom of an element in a molecule or an ion.*

Oxidizing agent *is the specie that performs the oxidation by taking electrons from the substance being oxidized.*

Periodic law *states that properties of the elements are the periodic function of their atomic numbers.*

Periodic Table *is arrangement of elements in order of their increasing atomic numbers in such a way that those with similar properties were placed one above other in the form of a table.*

Periods *are the horizontal rows in the periodic table.*

Polyatomic *molecule consists of two or more atoms.*

Radius of atom *or* **atomic radius** *is half of the distance between two atoms in an element measured using appropriate techniques.*

Reducing agent *is the specie that performs reduction by giving electrons to the substance being reduced.*

Reduction *is gain of electrons by an element or by a cation.*

Relative atomic mass *of an element is defined as the number of times one atom of that element is heavier than*

1/12(one-twelfth) the mass of one atom of carbon 12.

***Saturated solution** contains maximum amount of solute at a given temperature.*

***Shell** is the energy level in which electrons revolve around the nucleus e.g. K,L,M etc.*

***Shielding effect** is the decrease in attractive force exerted by the nucleus on the valence shell electrons due to the presence of electrons lying between the nucleus and the valence shell.*

***Solubility** is the number of grams of the solute dissolved in 100g of solvent to prepare a saturated solution at a particular temperature.*

***Solute** is the component of a solution, which is present in smaller quantity.*

***Solvent** is the component of a solution which is present in larger quantity.*

***Standard atmospheric pressure** is the pressure exerted by atmosphere at sea level.*

***Substance** is a pure piece of matter*

***Suspension** is a heterogeneous mixture of undissolved particles in a given medium. Particles are big enough to be seen with naked eye.*

***Unsaturated solution** is a solution which contains less amount of solute than that is required to saturate it at a given temperature.*

***Valence electrons** are electrons which are present in the outermost shell of an electron.*

INDEX

LIST OF BOOKS

1. Chemistry: The Central Science; T.L. Brown, H.E. LeMay, B.E. Burston Prentice Hall International, Inc.
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7. Understanding Chemistry: For Advanced Level; T Lister, J. Renshaw.
8. Chemistry Expression Science: EPB Pan Pacific.
9. Advanced Chemistry: Physical and Industrial; Cambridge.
10. New School Chemistry; O.Y. Ababio